DOCUMENT RESUME

ED 260 929 SE 045 969

AUTHOR Renner, John W.; And Others
TITLE Chemistry. Teacher's Guide. Investigations in Natural Science.
INSTITUTION Costa Rica Univ., San Jose. Research Inst. for the Improvement of Costa Rican Education.; Norman Independent School District 29, Okla.
PUB DATE 85
NOTE 297p.; For student investigations and readings see SE 045 970. Document contains light type. A joint project of the Science Education Center, University of Oklahoma and the Science Department, Senior High School of Norman Public Schools.
PUB TYPE Guides - Classroom Use - Guides (For Teachers) (052)

EDRS PRICE MF01/PC12 Plus Postage.
DESCRIPTORS *Chemistry; High Schools; *Learning Processes; *Science Activities; Science Education; *Science Experiments; *Science Instruction; *Secondary School Science; Teaching Methods

ABSTRACT

Investigations in Natural Science is a program in secondary school biology, chemistry, and physics based upon the description of science as a quest for knowledge, not the knowledge itself. This teaching guide is designed for use with the 19 chemistry investigations found in the student manual. These investigations focus on concepts related to: interactions with water; salt and calcium; physical properties; heating and cooling a pure substance; composition of a compound; atomic theory and matter; lead nitrate and potassium iodide; zinc combinations; hydrogen peroxide and manganese dioxide; mass relationships; temperature changes and dissolving; properties of ionic substances; Arrhenius solutions; pH, acids, and bases; acid-base interaction; salts and water; characteristics of methanol; solubility and structure; and periodicity and reactivity. Each investigation includes activities related to the three learning cycle components (exploration, conceptual invention, and expansion of the idea). The guide contains: a list of chemistry concepts fostered in each investigation; detailed instructional strategies for student investigations, which are reproduced in the teacher's manual in reduced form; student reading questions and answers (the readings, however, appear only in the student manual); and lists of equipment needed to complete experiments. (JN)

****
* Reproductions supplied by EDRS are the best that can be made *
* from the original document. *
****

ED260929

INVESTIGATIONS IN NATURAL SCIENCE

U.S. DEPARTMENT OF EDUCATION
NATIONAL INSTITUTE OF EDUCATION
EDUCATIONAL RESOURCES INFORMATION CENTER (ERIC)

✓ This document has been reproduced as received from the person or organization originating it.

☐ Minor changes have been made to improve reproduction quality.

• Points of view or opinions stated in this document do not necessarily represent official NIE position or policy.

CHEMISTRY

Teacher's Guide

John W. Renner
William T. Fix
Linda Cole Atkinson
Michael J. Renner
Michael R. Abraham

"PERMISSION TO REPRODUCE THIS MATERIAL HAS BEEN GRANTED BY

John W. Renner

TO THE EDUCATIONAL RESOURCES INFORMATION CENTER (ERIC)."

Science Education Center
College of Education
University of Oklahoma

and

Science Department
Senior High School
Norman Public Schools
Norman, Oklahoma

SE 045 969

Copyright 1985 by John W. Renner, William T. Fix,
Linda Cole Atkinson, Michael J. Renner, and
Michael R. Abraham

FOREWORD

Investigations in Natural Science is a program in secondary school biology, chemistry and physics based upon the description of science which says "Science is the quest for knowledge, not the knowledge itself."* That description of science is harmonious with that given by Einstein: "The object of all science is to coordinate our experiences and bring them into a logical system."** Notice that Einstein describes science as bringing our experiences into a system and does not describe science as the resulting system. Einstein's developing a system and Roller's quest for knowledge are describing the same thing; science is the search and the developing.

We asked ourselves what should students search or quest for in the three disciplines, and the obvious answer seemed to be understanding of the concepts of the disciplies. In order to state the concepts we believed students would develop during a quest, a detailed study was made of the intellectual capabilities of the students who enroll in the courses. Those findings were used in stating concepts which are given in each of the three courses. In determining the intellectual capabilities of the students the quality-of-thought model of Jean Piaget was used for guidance. In using that model, the recent modifications made in it by persons other than Piaget were considered.

An investigation to be conducted by the students which would lead them to each concept was prepared. These investigations--18 in biology, 18 in chemistry and 36 in physics--constitute what the students study in those three courses. A basic axiom of teaching was used in constructing the investigations. That axiom is: One person cannot teach another unless the teacher understands a model of how the learner learns. The learning model used was the mental functioning design of Jean Piaget. That model states that through interacting with the phenomenon (or concept) to be learned--and its associated materials--students assimilate the essence of the concept and that assimilation causes them to wonder, or puts the students in disequilibrium. The

---

*Duane H.D. Roller, "Has Science A Climate?", Sunday Oklahoman, Oklahoma City, Okla., February 22, 1970, p. 23.

**Gerald Holton and Duane H.D. Roller, Foundations of Modern Physical Science, Addison-Wesley Publishing Co., Inc., Reading, Mass., 1958, p. 214.

nature of a disequilibrated state leads students to re-equilibriate themselves by accommodating their mental structures and thinking processes to the assimilation. In essence, students put their thoughts in accord with those things which give rise to the assimilation. The new thought -- says Piaget's mental-functioning model -- now must be organized with pre-existing thoughts and ideas, or thought must be put in accord with thought. In this way the students construct their own knowledge of the concept investigated. These new constructs incorporate both the processes of science as well as its content and become a part of the students' mental structures rather than just a short-term set of memorized facts about the concept.

The mental-functioning model was used as the model of learning upon which the teaching plan in the Investigations in Natural Science was based. The assimilation experiences are provided by a series of activities under the heading Exploration. In those activities the students gather data which they and the teacher use to arrive at an accommodation which is the concept to be learned. That accommodation begins in the exploration but reaches fruition in a section entitled Conceptual Invention or The Idea. Organization experiences are included in a section called Expansion of the Idea. Each investigation includes a rich and sometimes a long expansion section. Teachers must use their judgment in deciding whether or not all the activities in each expansion will be used. We feel that the students may not always complete an entire Expansion section.

One part of the Expansion section of nearly every investigation which we believe is imperative is the student readings. Students read about a concept after they have experienced it in those programs. That is truly an expansion-of-the-idea activity. Reading about a concept before experiencing it -- the "advanced reading assignment" -- causes students to read about something they do not understand and that often leads to attempting to remember rather than understand what was read. The placement of the readings was carefully planned and we believe those readings serve their purpose. Teachers are urged to discuss the readings with their students; that discussion is essential to the Expansion-of-the-Idea phase of each investigation.

The three phase teaching plan and curriculum organization procedure used in constructing the investigation in all three disciplines -- Exploration, Conceptual Invention (The Idea) and Expansion of the Idea -- is known as the learning cycle. We cannot take credit for inventing the learning cycle; that credit goes to Dr. Robert Karplus, University of California, Berkeley, The Science

Curriculum Improvement Study and, of course, Dr. Jean Piaget.

Norman, Oklahoma
October, 1984

John W. Renner
Michael R. Abraham
Larry Atkinson
Linda Cole Atkinson
Jean McGregor Cate
William T. Fix
Eileen Bross Grzybowski
Edmund A. Marek
Jim A. Nickel
Michael J. Renner
Carl Surber
Susan Love Westbrook

ABOUT THE AUTHORS . . .

The authors of the INVESTIGATIONS IN NATURAL SCIENCE: CHEMISTRY program have two common places of employment--the Norman, Oklahoma, Public Schools and the University of Oklahoma. Those authors also have a second common characteristic; they are all knowledgeable about the Piagetian developmental-learning model and are interested in using it as a curriculum-construction model and a classroom-teaching procedure model.

John W. Renner is a Professor of Science Education and Adjunct Professor of Physics at the University of Oklahoma. Michael R. Abraham is an Associate Professor of Chemistry and Adjunct Professor of Science Education at that institution. William Fix is a teacher of chemistry and physics in the Norman Public Schools. Linda Atkinson was a chemistry teacher in the Norman Schools when the project began; she then moved to the science coordinator's position. She is presently science chairman for the middle schools and an assistant principal. Michael J. Renner also taught physics and worked on the chemistry investigations at one point in the programs' development and left to enter graduate school in psychobiology at the University of California, Berkeley. He is now on the psychology faculty at the University of Wyoming.

TABLE OF CONTENTS

Foreword		i
About the Authors		iv
Chemistry Concepts		vi
Chemistry Equipment and Materials List		viii
INVESTIGATION 1:	Interactions with Water	1
INVESTIGATION 2:	Salt and Calcium	13
INVESTIGATION 3:	Physical Properties	25
INVESTIGATION 4:	Heating and Cooling a Pure Substance	35
INVESTIGATION 5:	Composition of a Compound	46
INVESTIGATION 6:	Atomic Theory and Matter	57
INVESTIGATION 7:	Lead Nitrate and Potassium Iodide	112
INVESTIGATION 8:	Combinations with Zinc	120
INVESTIGATION 9:	Hydrogen Peroxide and Manganese Dioxide	132
INVESTIGATION 10:	Mass Relationships	143
INVESTIGATION 11:	Temperature Changes and Dissolving	164
INVESTIGATION 12:	Properties of Ionic Substances	176
INVESTIGATION 13:	Arrhenius Solutions	195
INVESTIGATION 14:	pH * Acids and Bases	205
INVESTIGATION 15:	Acid-Base Interaction	214
INVESTIGATION 16:	Salts and Water	221
INVESTIGATION 17:	Characteristics of Methanol	227
INVESTIGATION 18:	Solubility and Structure	259
INVESTIGATION 19:	Periodicity and Reactivity	269

CHEMISTRY CONCEPTS

Concept	Investigation Title
1. Solutions are homogeneous mixtures of two or more substances. Solutions are clear.	Interactions with Water
2. The changes in properties of a material can be either chemical or physical in nature. A chemical change cannot be reversed by physical means.	Water and Its Properties
3. Density is a measure of how tightly packed matter is or the amount of matter per unit volume. Density if a characteristic property.	Physical Properties
4. Melting point is the temperature at which a solid changes to a liquid, or a liquid changes to a solid, while the temperature remains constant. Boiling point is the temperature at which a liquid changes to a gas, or a gas changes to a liquid while the temperature remains constant.	Heating and Cooling a Pure Substance
5. Compounds have definite proportions of weight and numbers of atoms.	Composition of a Compound
6. An atom is the smallest particle of an element composed of protons, neutrons and electrons. The electrons are arranged in such a way that chemical properties of elements can be predicted from those arrangements.	Atomic Structure and Nomenclature of Common Ions
7. Chemical reactions involve changes in the combinations of atoms.	Lead Nitrate and Potassium Iodide
8. Some chemical reactions involve changes in the electron structure of atoms.	Combinations with Zinc
9. Catalysts are substances that speed up a chemical reaction without undergoing a net change.	Hydrogen Peroxide and Manganese Dioxide

	Concept	Investigation Title
10.	Chemical substances react in definite molar weight proportions.	Mass Relationships
11.	Reactions involve changes in the total energy of the reacting substances. The amount of energy is proportional to the amount of reacting substances.	Temperature Changes and Dissolving
12.	Soluble ionic substances dissolve because water interacts with the ions of the solid causing the ions to separate.	Properties and Ionic Substances
13.	Arrhenius acids are compounds that release H^+ ions in water solution; bases are compounds that release OH^- ions in water solution.	Arrhenius Solutions
14.	The molar concentration of the H ion determines the extent to which a solution is acidic or basic and is measured in units called pH.	pH
15.	Acids and bases can interact with each other to produce a neutral solution.	Acid/Base Interaction
16.	Compounds that donate H^+ ions in a reaction are acids, and compounds that accept H^- ions in a reaction are bases according to definitions advanced by Bronstead and Lowry.	Salts and Water
17.	Covalent bonding is a chemical bonding created by the mutual attraction for electrons between atoms in an element compound.	Characteristics of Methanol
18.	Characteristic physical properties of substances are related to the molecular or ionic structures of these substances.	Solubility and Structure
19.	The chemical properties (reactivities) of metals are determined by their electron structures.	Periodicity and Reactivity

CHEMISTRY EQUIPMENT AND MATERIALS LIST

AMOUNT	ITEM	INVESTIGATION NUMBER
1	asbestos pad	12
1	balance, laboratory (0.01 sensitivity)	2,3,5,7,10
5	balloons, round (per demonstration)	18
9	beakers, 50 ml	1,7,14,15, 16,17,20
1	beaker, 100 ml	2,7,16,20
4	beakers, 150 ml	2,11,13,18, 20
9	beakers, 250 ml	1,7,11,17, 18
4	beakers, 400 ml	4,7,8,10, 14,18,20
2	beakers, 600 ml	2,14
sev	boiling chips	4
1	burette clamp	10
1	burner, laboratory (Bunsen)	2,4,5,12, 13,20
1	candle	12
1	clamp for eudiometer	14
2	clamps, utility	4,10
10	corks to fit 13 x 100 mm test tubes	18,20
1	crucible cover	5
1	crucible tongs	5
1	D.C. ammeter	8
1	D.C. voltmeter (low range)	8
1	deflagrating spoon with wire screen to cover bowl	20
1	dropper pipette	16
30	dropper bottles	(7),8,15
1	dry cell, 1.5 v	8
1	electroscope	6
1	Erlenmeyer flask, 125 ml	10
10	Erlenmeyer flasks, 1000 ml (per class)	15,16
1	eudiometer, 50 ml	10, (14)
3	filter paper	7,11
1	Florence flask, 500 ml with stopper	13
1	forceps (for demonstration)	20
1	funnel, glass	7,11
1 pr	glass and rubber excitation rods	6,18
1	evaporating dish	5
1	glass stirring rod	1,7,11,13
1	glass tubing, 90 deg bend	10
1	glass tubing, U-bend	14

2	graduated cylinder, 10 ml	3,4,11,14
1	graduated cylinder, 25 ml	3,10,11
1	graduated cylinder, 50 ml	18
1	graduated cylinder, 100 ml	3
1	knife (for demonstration)	20
1	light bulb conductivity apparatus	1,2,13,14 16,17,18
1 set	litmus paper, red and blue	1,13,14, 17,18
4	magnets, disc (per demonstration)	18
1	magnetic stirrer with magnet bar (per class)	16
sev	matches	12,20
1 pr	metal and glass rods for heating experiment	(6)
1 roll	paper towels	20
1 set	pH paper (per class demonstration)	17
2	pipettes, 10 ml with suction bulbs	15,16
1	pipestem triangle	5
2	pith balls on strings	6
4	plastic spoons	1,2,10,12, 13,14,17,19
1	plastic weighing cup	7
2	ring stands	2,4,5,6,7, 10,11,13,14, 16,18,20
2	ring stand rings, large	2,4,6,20
1	ring stand ring, small iron	5,7,11
1 m	rubber tubing	10,14
1	scissors	5
1 ea	silk and wool cloth	6,18
some	steel wool, fine	5,9
1	stopper, one-hole rubber number 4	10,14
1	stopper to fit 18 x 150 test tube	3,20
2	stoppers to fit 10 ml graduated cylinder	3
4	styrofoam cups, 300 ml	12
12	test tubes, 13 x 100 mm	(7),11,19,20
4	test tubes 18 x 150 mm	2,3,4,5,8,9, 12,20
1	test tube, 25 x 150 mm	14
1	test tube clamp	5,12
1	test tube rack	(7),8,9,11, 12,18,20
1	thermometer, -10 to 110 C	4,12
3	wash bottles, 250 ml	7,18
2	watch glasses	2,20
1	wire gauze with asbestos center	2,4,[illegible]
6	wires with alligator clips	8
sev	woodsplints	4,10

INVESTIGATION 1: INTERACTIONS WITH WATER

TEACHER'S GUIDE

INTRODUCTION

The Exploration phase of this investigation provides activities which lead to the conceptual invention of solution. As the solubility characteristics of the various substances with water are disussed, additional terms will need to be invented. These necessary terms are: soluble, insoluble, solvent, solute, miscible, and immiscible. Conductivity of solutions and acid/base indications with litmus paper should only be treated as observable data. The discussions involving the flow of ions during conductivity and a solution's reaction with litmus paper will be held during Investigations 13 and 14, respectively.

CONCEPT

Solutions are homogeneous mixtures of two or more substances. Solutions are clear.

EQUIPMENT

A. Exploration

1. Labware

a. light bulb conductivity apparatus
b. litmus paper (red and blue)
c. 250 ml beakers (steps 1-2)
d. 50 ml beakers (steps 3-5)
e. glass stirring rod
f. plastic spoon

2. Reagents

a. distilled water
b. for steps 3-5, select 7 of the following:

(1) iodine (solid)
(2) hexane
(3) ethanol
(4) iron (III) chloride (solid)
(5) cyclohexane
(6) oleic acid
(7) charcoal powder (solid)
(8) potassium chloride (solid)
(9) magnesium oxide (solid)
(10) calcium oxide (solid)

INVESTIGATION 1: INTERACTIONS WITH WATER

NOTE: Iron (III) chloride and magnesium oxide will form acidic and basic solutions, respectively. For this reason, each of them should be used. Also, it is helpful in Investigation 5 for the students to have tested magnesium oxide.

B. Expansion of the Idea

1. Labware

a. light bulb conductivity apparatus
b. litmus paper (red and blue)
c. 250 ml beakers

2. Reagents

a. aluminum chloride
b. potassium iodide
c. copper (II) sulfate
d. calcium sulfate
e. iodine in carbon tetrachloride and water
f. sucrose
g. sodium carbonate
h. magnesium chloride
i. sodium phosphate
j. carbon tetrachloride
k. distilled water

INVESTIGATION 1: INTERACTIONS WITH WATER

To prepare these solutions:

(a) $AlCl_3$: 0.8 g/100 ml H_2O

(b) KI: 1.0 g/100 ml H_2O

(c) $CuSO_4$:

4.0 g $CuSO_4 \cdot 5H_2O$/100 ml H_2O

(d) $CaSO_4$: 0.75 g/100 ml H_2O

(Heat to at least 60C, then filter.)

(e) I_2: 0.3 g I_2/200 ml CCl_4 and H_2O

(f) $C_{12}H_{22}O_{11}$: Any concentration solution.

(g) Na_2CO_3: 1.0 g/100 ml H_2O

(h) $MgCl_2$: 1.0 g/100 ml H_2O

(i) Na_3PO_4: 4 g/100 ml H_2O

INVESTIGATION 1: INTERACTIONS WITH WATER

PRE-LAB INSTRUCTIONS

The procedures in this investigation--use of the conductivity apparatus, reading liquid levels in a graduated cylinder, and use of litmus paper--need a thorough discussion during the pre-lab session.

A. Conductivity Apparatus

1. Be sure to stress that the ONLY TIME THE APPARATUS SHOULD BE PLUGGED IN IS DURING THE ACTUAL TESTING OF A LIQUID FOR CONDUCTIVITY.

2. The apparatus is most easily manipulated if it is clamped on a ring stand and then raised and lowered by loosening the holding bolt.

3. The electrodes should be rinsed with distilled water before being transferred from one solution to another.

4. A word about the disciplinary and possible physical hazards associated with irresponsible use of the conductivity apparatus and wash bottles is needed here.

B. Use of Graduated Cylinders

1. Water volumes in this investigation are approximate. However, this is a good time to instruct the students in the proper method for reading liquid levels in graduated cylinders.

2. Demonstrate the method for obtaining a precise amount of water (using a wash bottle).

3. Invent <u>meniscus</u> at the appropriate time in this discussion.

C. Litmus Paper

1. A discussion of the use of litmus paper is needed. It should be limited to the visual interpretation that red litmus paper turning blue indicates the presence of a basic solution and blue litmus paper turning red indicates the presence of an acidic solution.

INVESTIGATION 1: INTERACTIONS WITH WATER

EXPLORATION

INVESTIGATION 1: INTERACTIONS WITH WATER

Name_____

Section_____

INTRODUCTION

"Interaction" denotes the effect that two or more substances have on each other when they are combined. You have probably observed that oil floats on water, that dirt is carried along by water in rivers, and that fruit punch soft drink powder causes water to be a reddish-orange color. Although you will not be using oil, dirt and soft drink powder in the laboratory, some of the interactions that you will perform will resemble these interactions. In what ways are the interactions of oil, dirt and soft drink powder with water similar to each other? In what ways are they different? How would you try to convince someone that your stated similarities and differences were reasonable?

In this investigation you will combine water and one other substance in several separate experiments. Your tasks will be to determine the extent to which various substances interact with water and to determine some basic similarities and differences in the interactions of substances with water.

EXPLORATION

1. Put approximately 100 ml of distilled water in a clean, dry 250 ml beaker. Dip one piece of blue litmus paper, then one piece of red litmus paper in the water and record your observations.

2. Place the conductivity apparatus in the beaker, according to your teacher's instructions. Plug in the apparatus and record your observations.

NOTE: Save the distilled water to use in the remainder of the tests.

INTRODUCTION: Allow the students to propose answers to the questions. Be careful not to make the invention at this point.

2. CAUTION: THE ONLY TIME THAT THE CONDUCTIVITY APPARATUS SHOULD BE PLUGGED IN IS DURING THE ACTUAL TESTING OF A LIQUID FOR CONDUCTIVITY.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

INVESTIGATION 1: INTERACTIONS WITH WATER

3. Your teacher will provide you with various liquid and solid substances. For each test use 25 ml of water and 3 ml of liquid (or a pea-sized amount of solid) in a 50 ml beaker.

4. For each test describe three things:

   a. What happens to the substance immediately when it is placed in the 50 ml beaker with the water?

   b. What happens to the substance after it has been stirred vigorously?

   c. Describe what effect the mixture has on litmus paper and whether the mixture will conduct electricity.

5. Record your observations in the table below.

Substance (Name and Formula)	Before Stirring	After Stirring	Litmus/ Conductivity

6. When the tests are completed, dispose of the mixtures according to your teacher's instructions.

3. Each liquid and solid substance should be in a covered container labeled with the name and formula of the substance.

6. DISPOSAL INSTRUCTIONS: Instruct the students to wash the inorganic substances down the drain. The organic wastes should be placed in a labeled bottle.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

CONCEPTUAL INVENTION (THE IDEA)

INVESTIGATION 1: INTERACTIONS WITH WATER

THE IDEA

1. Develop a classification system to describe various ways that substances interact with water.

2. What evidence can be used to show that some of these materials interact (mix) with water?

3. Considering the systems in which the water interacted with the other substance, make a statement describing the extent to which the substance mixes with the water in different regions of the mixture.

1. Disappears or dissolves, slightly dissolves, or makes the water cloudy.

2. The substances disappear. The light bulb lights. There is a color change in the litmus paper.

3. There are two possibilities:

   a. The substances mixed evenly in all regions.
   b. The substance is more concentrated at the bottom than at the top.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

INVENTION DISCUSSION

A. The main purpose of this discussion is to arrive at some agreement on the observations made during Exploration. In the process of doing this, the teacher will want to invent terms as they can be brought into the discussion. The terms to be invented are soluble, insoluble, miscible, immiscible, solvent, and solute.

B. Start the discussion with the observations made when distilled water was tested with the conductivity apparatus and litmus paper.

C. Have the students describe what happened to the various substances when placed in water. The following are questions which will elicit observational responses:

1. What happened when the substances were placed in distilled water?

2. Which of the substances disappear when placed in the water?

3. If the solid disappeared, what happened to it? (Dissolved--broke apart.)

4. Which of the substances dissolve?

5. Was there a color change when any of the substances were placed in water? If so, what color changes?

6. Did all of the substances mix with water?

7. If a substance did not mix, what observations did you make? (couldn't see through them or there was a line dividing the two liquids)

8. Look at the secondary evidences (i.e., litmus paper change and conductivity). Identify the substances which totally mixed that conduct an electric current and change the litmus paper. Also identify the substances which did not totally mix with water, but that conduct an electric current and change litmus paper. Does the change in litmus paper or conducting an electric current indicate the substance has partially mixed with water? (Conductivity and litmus tests can be performed on the solids, as a demonstration.)

BEST COPY AVAILABLE

9. Iron (III) chloride turns the liquid yellow. Does the liquid appear to be more yellow at the bottom, the middle, or the top? (It's evenly mixed throughout.)

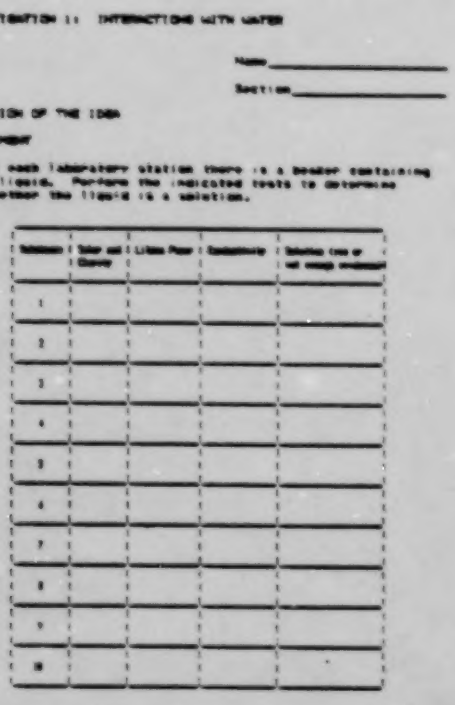
D. Invent <u>solution.</u> Solutions are homogeneous mixtures of two or more substances. Solutions are clear.

NOTE: Charcoal tends to color the mixture. Students may feel as if the charcoal has partially mixed with the water. Discuss the observations that charcoal and water do not change litmus paper or conduct an electric current. If the students are not satisfied that the charcoal is insoluble, have them take a sample and filter it through double filter paper.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

EXPANSION OF THE IDEA



INVESTIGATION 1: INTERACTIONS WITH WATER

Name_____

Section_____

EXPANSION OF THE IDEA

EXPERIMENT

1. At each laboratory station there is a beaker containing a liquid. Perform the indicated tests to determine whether the liquid is a solution.

Substance	Color and Clarity	Litmus Paper	Conductivity	Solution (yes or not enough evidence)
1				
2				
3				
4				
5				
6				
7				
8				
9				
10				

1. At each laboratory station place a 250 ml beaker containing 100 ml of one of the substances listed under 'Reagents' in the equipment section of the teacher's guide. The students are to rotate through the stations.

   Several tests can be used to identify whether a substance is a solution or not. The tests the students have used are conductivity, litmus paper, color, and clarity. For some substances this will not be enough evidence to determine whether it is a solution (for example, sucrose in water).

1. The discussion should begin by reviewing the data and deciding if each of the substances is a solution.

2. Iodine and carbon tetrachloride solution may be a surprise. Iodine and carbon tetrachloride form a solution which is immiscible with water.

3. The students will probably suggest that sugar and water do not form a solution, since it will not conduct a current or change litmus paper.

4. During the discussion bring up the point that solutions can be colorless such as the sugar solution and potassium iodide solution. Of course the students would need to run further tests to determine that sugar and water form a solution. The discussion of further tests on the sugar solution should be deferred until after Reading a is completed.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

Assign READING a: THE PROBLEM WITH SOLUTIONS, with its questions. The questions and acceptable answers follow.

INVESTIGATION 1: INTERACTIONS WITH WATER

READING 1a: QUESTIONS

1. What characteristics of solubility have we used to identify a solution?

2. If the solubility of potassium chloride is 35 g per 100 ml of water, how many grams should you be able to dissolve using 5 ml of water?

3. The sucrose ($C_{12}H_{22}O_{11}$) solution and distilled water had no effect on litmus paper, did not conduct, or change color. If you were given a clear, colorless liquid, explain how you could determine if it was water or a sucrose solution. NOTE: Tasting is not an option.

1. Change in litmus paper, conductivity, color, clarity.

2. $$\frac{35\ g}{100\ ml} = \frac{Xg}{5\ ml}$$

$$Xg = 1.75\ g$$

3. Allow the liquid to evaporate. If there is a solution present, a residue of sucrose will remain. No residue will remain after evaporating pure water.

BEST COPY AVAILABLE

INVESTIGATION 1: INTERACTIONS WITH WATER

Assign READING b: HOW HOMOGENIZED IS HOMOGENOUS? and its questions. The questions and acceptable answers follow.

[illegible] INTERACTIONS WITH WATER

Name _____

Section _____

READING (b) QUESTIONS

1. A 15 ml sample of a solution contains 3.0 grams of solute. How much solute would be dissolved in: (Show your work.)

   a. 1 liter

   b. 100 ml

   c. 1 ml

2. Explain how air can be considered as a solution.

3. Explain how air can be considered as a colloidal dispersion (include at least two examples).

11

Acceptable Answers:

1. (a) $\frac{3\ g}{15\ ml} = \frac{X\ g}{1000\ ml}$

$X\ g = 200\ g$

(b) $\frac{3\ g}{15\ ml} = \frac{X\ g}{100\ ml}$

$X\ g = 20\ g$

(c) $\frac{3\ g}{15\ ml} = \frac{X\ g}{1\ ml}$

$X\ g = 0.2\ g$

2. Air is a homogeneous mixture of 2 or more gases.

3. Answers such as small dust particles, pollen, water vapor, and so forth, are acceptable.

The discussion of the readings should consist of going over the questions after the students' papers have been checked. Also, the teacher should go through the readings and choose the content he/she wants to emphasize.

BEST COPY AVAILABLE

INVESTIGATION 2: SALT AND CALCIUM

TEACHER'S GUIDE

INTRODUCTION

During this learning cycle physical and chemical change are investigated. The main difference to be stressed between a physical and chemical change is that a physical change is reversible by physical means. During the Exploration phase students investigate a physical system (salt and water) and a chemical system (calcium and water) simultaneously. This enables the students to make comparisons in the two types of changes in materials. Chemical change should be discussed from observational evidence only, then later in the course the different types of chemical reactions will be studied. During the discussion the terms electrolyte and precipitate should be introduced as well as the formulas for the diatomic elements.

CONCEPT

The changes in properties of a material can be either chemical or physical in nature. A chemical change cannot be reversed by physical means.

EQUIPMENT

A. Exploration

1. Labware

a. light bulb conductivity apparatus
b. ring stands with large rings
c. litmus paper (red and blue)
d. beakers-100 ml and 150 ml
e. laboratory burners
f. wire gauze with asbestos center
g. plastic spoons
h. watch glasses

2. Reagents

a. distilled water
b. sodium chloride-approximately 50 g
c. calcium metal chips

INVESTIGATION 2: SALT AND CALCIUM

B. Invention Discussion-Demonstration I

1. Labware

a. 600 ml beaker
b. 3-18x150 mm test tubes
c. wood splints

2. Reagents

a. tap water
b. calcium metal chips

C. Expansion of the Idea (Demonstration)

1. Labware

a. 2-600 ml beakers
b. 1-250 ml beaker

2. Reagents

a. ammonium hydroxide
b. copper (II) sulfate solution
c. aluminum pieces
d. hydrochloric acid-2M

INVESTIGATION 2: SALT AND CALCIUM

PRE-LAB INSTRUCTIONS

One class period will be required for the pre-lab activities and for the students to weigh their systems in the Exploration phase. Pre-weighed objects can be used for student practice prior to beginning the exploration for the investigation.

A. These procedures are written specifically for the single-pan, Dial-O-Gram balance, but they apply in general to most types of laboratory balances.

1. Set the balance on the demonstration table so that the class can see the graduations, and so that it balances on zero. Point out the position of the arm in the balanced condition. Go over the ranges of weight on the balance arms and the dial.

2. Tell the students to treat the balances carefully because of their delicacy. If they move them at all, they must do it very carefully to avoid getting them out of balance. If balance is disrupted, the teacher will make the balancing adjustments. Of course, the balance needs to be checked on zero before measurements are made.

3. Weigh some object by placing the object on the balance pan, moving the coarse weights to within 10g, and then moving the dial to the balance position.

4. Go through the reading of the weight step by step, using students to take the readings, if you wish. Write the weight on the board.

5. Take a second and third reading on the object by turning the dial back toward zero, moving the object to a different part of the pan, and then dialing to rebalance. Record these weights on the board and illustrate finding the average value of the readings.

6. This demonstration should point out the reason that the students are asked to weigh the various systems three times instead of once.

B. The use of the laboratory burner is also a new procedure in this investigation. There are some procedures and safety instructions that should be given to the students before they begin working with their burners.

INVESTIGATION 2: SALT AND CALCIUM

1. The parts of the burner and their function should be shown to the students; for example, the device on the bottom of the burner adjusts the amount of air supplied to the burner.

2. Tell the students that the match should <u>always</u> be burning <u>before</u> the gas is turned on. Show the class how to set the burning match at the side of the burner tube to prevent a strong stream of gas from blowing out the flame.

3. Students should be advised to check to be sure when the air lever is fully closed and fully open. Some levers become loose and fail to seal properly. It is also possible to push the lever so far in the "closed" direction that it actually opens again a small amount.

4. Perform a demonstration of the use of the laboratory burner. Light the burner and ask the students to observe the flame. Have the students compare the flame when the air vent is closed to the flame when the air vent is open. Hold a piece of asbestos paper with crucible tongs in the flame of the burner. The students should decide which part of the burner flame is the hottest. Perform this procedure with the air vent closed and then with the air vent open. (The hottest point of the flame is just above the top of the inner cone.)

INVESTIGATION 2: SALT AND CALCIUM

EXPLORATION

INVESTIGATION 2: SALT AND CALCIUM

Name_____

Section_____

INTRODUCTION

This investigation will involve the interaction of two more substances with water. One major difference between this and Investigation 1 is that you will collect weight data on the solid substances before adding water and at the end of the exploration procedure. How would you expect the weight of a solid substance to change if water were added to it to make a solution and then the water were evaporated to leave a solid substance?

EXPLORATION

1. Clean and dry your 100 ml and 150 ml beakers and two watch glasses. Weigh each of them three (3) times. Record the weights to two (2) decimal places.

100 ml beaker and watch glass	150 ml beaker and watch glass
_____	_____
_____	_____
_____	_____
average_____	average_____

NOTE: Round your average weight to two decimal places. Always keep the same beaker and watch glass combinations.

2. In one of the beakers place a spoonful of salt (sodium chloride—NaCl), and in the other beaker place a designated amount of calcium metal (Ca).

18

3. Weigh each beaker, watch glass, and its contents three times and record the weights to two decimal places.

100 ml system and NaCl	150 ml system and Ca
_____ g	_____ g
_____ g	_____ g
_____ g	_____ g
average_____ g	average_____ g

4. From your data, determine the weights of salt and calcium contained in the system. Show your calculations.

Salt

Calcium

5. Add 25 ml of distilled water to the beaker containing salt. Replace the watch glass and record careful observations of the interaction between salt and water.

6. After the observations are written, test the remaining solution for conductivity and with litmus paper. Record the results of these tests.

7. Transfer the salt solution to a ring stand. Heat the system strongly until it is completely dry. While this system is heating, you may begin working with the calcium system.

18

17

In Step 3, and for the remainder of the laboratory, the watch glass should be placed on the beaker concave side up.

Spattering of the solid should be carefully avoided during the heating process. It may be necessary to remove the burner from time to time in order to prevent spattering (Steps 7 and 18).

BEST COPY AVAILABLE

INVESTIGATION 2: SALT AND CALCIUM

INVESTIGATION 2: SALT AND CALCIUM

9. Add 25 ml of distilled water to the beaker containing calcium. Replace the watch glass and record careful observations of the interaction between calcium and water.

9. Test the remaining liquid for conductivity and with litmus paper. Record the results of the tests.

10. When the salt system is completely dry, remove it from the ring stand and place the calcium system on the ring stand. Heat the calcium system strongly until it is completely dry.

11. After the dried systems cool to room temperature, weigh each system and record the weights.

salt system	calcium system
_____ g	_____ g
_____ g	_____ g
_____ g	_____ g
average _____ g	average _____ g

12. Determine the weights of the substances remaining in the two beakers. Show your calculations.

14

Prior to weighing the systems in Step 11, allow the systems to cool to the extent that the beakers can be held comfortably in the hand. Caution the students that glass cools slowly.

Instruct the students to put their beakers in the drawer after the final weight has been determined.

In Steps 13 and 14, the students are to calculate percentage weight change. This process will be facilitated by a discussion of the formula used to find the percentage weight change.

% weight change =

$$\frac{\text{final weight} - \text{initial weight}}{\text{initial weight}} \times 100$$

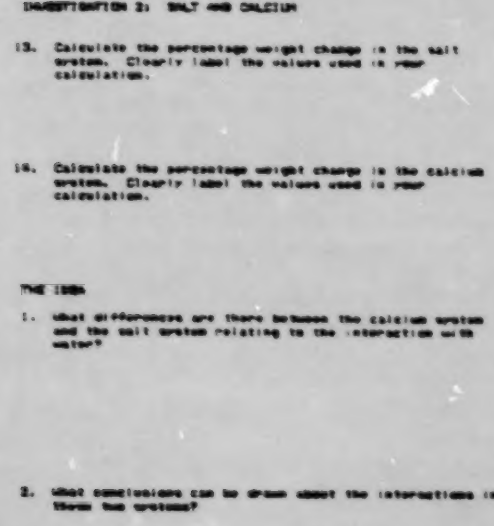
Stress to the students that the terms final weight and initial weight are referring to the weight of the <u>substance only.</u> These weights do not include the weight of the beaker and watch glass. Also, it will be helpful to determine, through a brief discussion, that the weights in Step 12 are the final weights and the weights in Step 4 are the initial weights.

BEST COPY AVAILABLE

BEST COPY AVAILABLE

INVESTIGATION 2: SALT AND CALCIUM

CONCEPTUAL INVENTION (THE IDEA)



INVESTIGATION 2: SALT AND CALCIUM

13. Calculate the percentage weight change in the salt system. Clearly label the values used in your calculation.

14. Calculate the percentage weight change in the calcium system. Clearly label the values used in your calculation.

THE IDEA

1. What differences are there between the calcium system and the salt system relating to the interaction with water?

2. What conclusions can be drawn about the interactions in these two systems?

15

Acceptable Answers:

1. Suggest that the students list the characteristics of the two systems and compare the list to determine the differences in the two systems.

2. The interactions are different.

INVENTION DISCUSSION

A. Ask the class about the percentage weight change in their salt system and calcium system (that is, did the weights increase, decrease, or remain the same).

   1. When a consensus has been achieved, have each group of students list the percentage weight change data for the salt and calcium systems on the board or overhead transparency.

   2. Ask the class to compute an average value for the percentage weight change in the salt system and the calcium system. Have the students compare these values.

B. Have the students describe their observations of the salt and water system and then the calcium and water system. List these observations on the board.

BEST COPY AVAILABLE

C. One of the observations cited should be that there were bubbles produced when the water was added to the calcium. Pursue this observations by asking what the bubbles imply (gas production). Lead the class to a general statement formulating the idea that the presence of bubbles indicates the formation of a gas. When the class is convinced that a gas is produced, proceed with the following demonstration.

NOTE: The apparatus required will be a beaker (or clear plastic tray) half filled with water, three 18x125 mm pyrex test tubes, three chips of calcium metal, and one or two wood splints.

1. Explain that one test that is usually performed on gases is a test for flammability.

2. Solicit a volunteer from the class to manipulate the wood splint. Place a piece of calcium in the water, and hold an inverted test tube (full of water) over the piece of calcium and collect a test tube full of the hydrogen evolved.

NOTE: If the container used is not sufficiently large to permit direct filling of the test tube, there is an easy method for obtaining a test tube full of water. Fill the test tube with water and then force a solid rubber stopper into the test tube. Invert the stoppered test tube in the beaker (or tray) and remove the stopper with foreps or crucible tongs. Since tap water is acceptable for this demonstration, you may want to fill the test tubes, by this procedure, while the class is watching. This procedure may be of use to them in later investigations.

3. When you have collected a test tube full of the gas, have the student volunteer light the splint and hold it (while burning) to the mouth of the test tube, while you hold the test tube sideways. Ask the students to interpret their observations. You may be pressed, by popular demand, to repeat the above procedure.

4. Collect hydrogen gas in a test tube full of air and compare this flammability test to the test made on the test tube full of hydrogen that was collected over water.

5. Fill another test tube with gas, but this time hold the test tube upright for about 5 seconds before having your assistant test it with a burning splint. Ask for an interpretation of this experiment.

6. Name the gas (hydrogen) and briefly review the characteristics that have been observed.

D. Invent the formula for the gas, hydrogen (H_2). Tell the students that the symbol for hydrogen, when written alone, must <u>always</u> have the subscript <u>two</u> included. If the symbol H_2 does not seem reasonable to the students, ask them to accept the formulas for the time being. The reasons for writing the formulas in this way will be explored later.

Tell the students that there are seven elements that must be written with the subscript two <u>when they exist alone</u> (emphasize this point). They are hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine. Write these elements' names on the board along with their formulas. Tell the students that they should memorize this fact about these elements. All others <u>can be</u> written, when alone, using only the symbol for the element. As an aid in the memorization of these special elements, point out their locations on the periodic chart.

E. Return the discussion to the observations of the salt and calcium experiments. Although there appeared to be a dissolving process in both systems, there were significant differences in the two systems. Ask students to identify these differences from the list of observations.

F. Ask the students what was left in the beaker from the salt and water system. Have a volunteer return to their laboratory station and add the water to their salt beaker. Did the substance in the beaker behave the same as the salt behaved when water was added? (Yes, it dissolved.)

G. Follow the same line of questioning with the calcium system. "What was in the calcium beaker after the experiment?" "What could we do to test the substance in the beaker to see if it is calcium metal?" (Add water and see if gas is produced.) Send the class to their laboratory stations to clean the beakers from their salt and calcium systems. Discuss the observations. "What do your weight data indicate about the substance in the beaker?" (It weighs more than the calcium metal.)

H. Using the observational differences for the salt and calcium systems, invent the concepts of <u>physical change</u> and <u>chemical change</u> (reaction).

Changes in properties of the material can either be chemical or physical in nature. The main difference between chemical and physical change is that a physical change is reversible by physical means. A chemical change is defined as a change in matter that cannot be reversed by physical means.

Define the salt and calcium system in terms of physical or chemical change. (The salt underwent a physical change, and the calcium underwent a chemical change.)

I. Ask the students if a solution formed when calcium reacted with water. There will be a difference of opinion. Some students will feel it did not form a solution because it was not clear. Others will point to the fact that it conducted and changed litmus paper.

DEMONSTRATION II

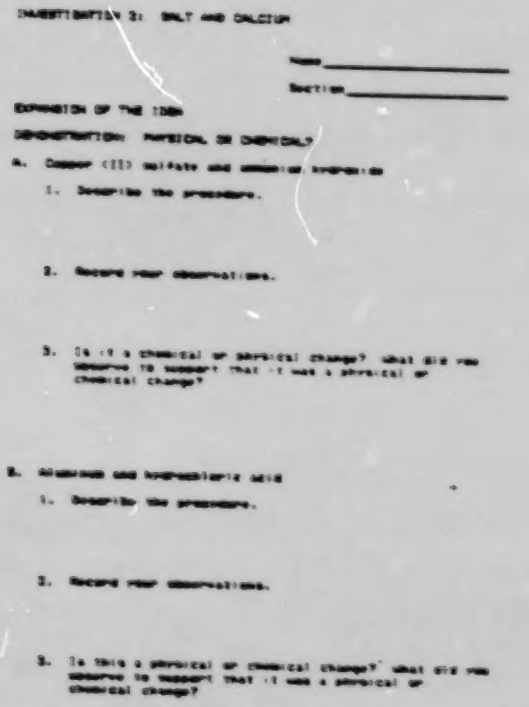
Put one or two chips of calcium metal (one chip at a time to avoid getting a precipitate) into a beaker containing 200-300 ml of water. Ask the students to observe carefully. Test with litmus paper and for conductivity (gas is produced, solution is clear, litmus paper changes from pink to blue, it may conduct). "What does this show?" (A solution was formed.) Put in several pieces of calcium metal. Let them react. Have the students observe carefully. (A precipitate is formed.)

"Was that white substance in there when we started?" (no), and "When did the white substance form?" are questions to ask the students. Discuss the formation of a precipitate.

J. Direct the attention of the students to obserations that indicate chemical change (for example, evolution of a gas, permanent change in physical properties, production of heat). Mention that one observation indicating a chemical change cannot always be conclusive evidence that a chemical change has taken place. For example, salt will remain in solution indefinitely in a closed system.

INVESTIGATION 2: SALT AND CALCIUM

EXPANSION OF THE IDEA



Name_____

Section_____

Review definitions of chemical change and physical change. Students should complete student sheet I2-5 during the demonstration.

Section A
- Pour copper (II) sulfate solution in a beaker.
- Have students make their observations of color, clarity, litmus paper, etc.
- Pour ammonium hydroxide in a beaker. Have students make their observations of odor, color, clarity, litmus paper, etc.
- Add the ammonium hydroxide solution to the copper (II) sulfate solution. (A deep purple color and a precipitate form.
- Discuss the observational changes. Was it a chemical or physical change? (Chemical.) Have students supply supporting information.

Section B
- Show the aluminum to the class.
- Give the name and formula for the hydrochloric acid.
- Add the aluminum to the HCl.
- Have the students write their observations.
- Discuss whether this is a chemical or physical change.
- List the evidence.
- Is it possible to have a gas produced in a physical change? (Yes. The boiling of a liquid substance is such an example.)

BEST COPY AVAILABLE

INVESTIGATION 2: SALT AND CALCIUM

Assign the READING : FOOD FOR THOUGHT AND BREAKFAST, and its questions. Follow with a discussion of the salient points in the reading.

INVESTIGATION 2: SALT AND CALCIUM

Name_____

Section_____

READING 2: QUESTIONS

For each of the following situations, state whether the process is chemical or physical, and then list supporting evidence.

1. Fermentation.

2. Leaves turning brown in the fall.

3. Running a gasoline engine.

4. Cooking a piece of bacon.

5. Popping popcorn.

6. Effervescent antacid tablet in water.

Acceptable Answers:

1. Chemical. Gas produced, properties, changed, new substances produced.

2. Chemical. Color change, texture change.

3. Chemical. State change, odor change, heat produced.

4. Chemical. Color change, texture change, odor change.

5. Physical. The evaporation of trapped moisture causes the explosion.

6. Chemical. Gas produced, litmus change, conductivity change.

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES

TEACHER'S GUIDE

INTRODUCTION

This investigation is, in some ways, out of the mainstream of content development in the course. The concept that is invented, however, lends itself to a better understanding of material throughout the course. There are points in the course (see Investigation 18) where a good basis in this idea is highly important.

CONCEPT

Density is a measure of how tightly packed matter is or the amount of matter per unit volume. Density is a characteristic property.

EQUIPMENT

A. Exploration

1. Labware

   a. laboratory balances (0.01g sensitivity)
   b. 18 x 150mm test tube
   c. stopper to fit 18 x 150mm test tube
   d. 25 ml graduated cylinder
   e. 10 ml graduated cylinder

2. Reagents

   a. lead shot
   b. mossy zinc
   c. aluminum shot
   d. tert-butyl alcohol
   e. tap water

B. Expansion of the Idea (Demonstration I)

1. Labware

   a. 2-10 ml graduated cylinders
   b. laboratory balance
   c. stoppers to fit graduated cylinders

2. Reagents

   a. lead nitrate solution
   b. distilled water

INVESTIGATION 3: PHYSICAL PROPERTIES

C. Expansion of the Idea (Demonstration II)

1. Labware

a. 2-100 ml graduated cylinders

2. Reagents

a. carbon tetrachloride
b. distilled water
c. hexane

INVESTIGATION 3: PHYSICAL PROPERTIES

EXPLORATION

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

INTRODUCTION

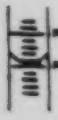
In this investigation you will measure the weights and volumes of three metals and one liquid. The volumes of the solids will be obtained by putting them in water and observing the volume of water displaced (or pushed out of the way). This is a convenient method for determining the volume of an irregularly shaped object. Why is it that the water displacement method cannot be used to find the volume of some substances?

EXPLORATION

1. Obtain samples of lead (Pb), zinc (Zn), and aluminum (Al). Compute the average weight of each sample and record the weights in the appropriate spaces below.

	Lead	Zinc	Aluminum
	_____	_____	_____
	_____	_____	_____
	_____	_____	_____
average	_____	_____	_____

2. Determine the volume of each sample separately as follows:



a. Fill a 25 ml graduated cylinder to the 10 ml mark. Liquid levels should be read from the bottom of the dip (or meniscus)—see the diagram at left.

b. Carefully add one of the samples of metal to the graduated cylinder. Be sure that no air bubbles remain trapped on the metal.

INVESTIGATION 3: PHYSICAL PROPERTIES

c. Find the sample volume by subtracting the (initial) liquid level from the final liquid level:

(water and sample in ml) − (water alone in ml) = (sample in ml)

d. Repeat this procedure for all three samples and record the sample volumes in the spaces below.

lead_____ml

zinc_____ml

aluminum_____ml

e. Put the used metal samples in the place designated by your teacher.

3. Clean and dry a 10 ml graduated cylinder. Weigh the graduated cylinder separately. Compute its average weight and record the weights in the appropriate spaces below.

_____

_____

_____

average _____

4. Your teacher will provide you with a sample of terrestrial element; which should be poured into your 10 ml graduated cylinder. Read the volume of this sample carefully and record it below.

terrestrial element_____ml

1. The metals should be put in containers that are accessible to students. Also, have return containers for substances after they are wet so they can be reused.

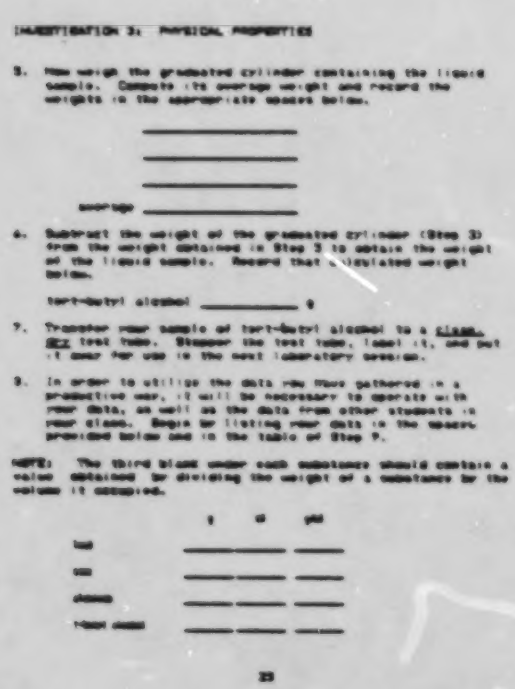
Remind the students to include units on the measurements they make.

2. Stress the need to get all the air bubbles out of the system. Ask the students why such care is important.

Stress that the 10 ml graduated cylinder (Step 3) and the test tube (Step 7) must be clean and dry. It is very important for the substance to be pure for use in Investigation 4.

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES



INVESTIGATION 3: PHYSICAL PROPERTIES

3. Now weigh the graduated cylinder containing the liquid sample. Compute its average weight and record the weights in the appropriate spaces below.

_____

_____

_____

average _____

4. Subtract the weight of the graduated cylinder (Step 3) from the weight obtained in Step 3 to obtain the weight of the liquid sample. Record that calculated weight below.

tert-butyl alcohol _____ g

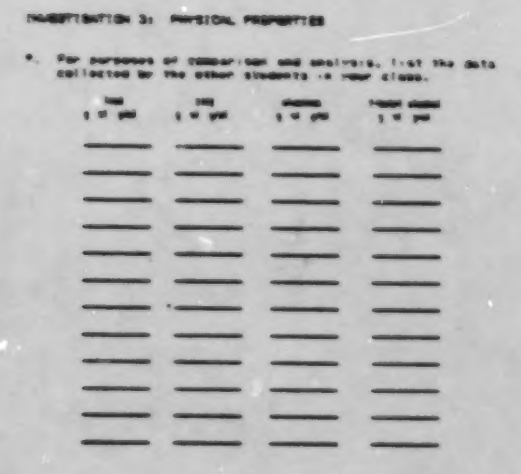
7. Transfer your sample of tert-butyl alcohol to a clean, dry test tube. Stopper the test tube, label it, and put it away for use in the next laboratory session.

8. In order to utilize the data you have gathered in a productive way, it will be necessary to operate with your data, as well as the data from other students in your class. Begin by listing your data in the spaces provided below and in the table of Step 9.

NOTE: The third blank under each substance should contain a value obtained by dividing the weight of a substance by the volume it occupied.

	g	ml	g/ml
[illegible]	_____	_____	_____
[illegible]	_____	_____	_____
[illegible]	_____	_____	_____
[illegible]	_____	_____	_____

25



INVESTIGATION 3: PHYSICAL PROPERTIES

9. For purposes of comparison and analysis, list the data collected by the other students in your class.

[illegible] g ml g/ml	[illegible] g ml g/ml	[illegible] g ml g/ml	[illegible] g ml g/ml
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

26

4. The teacher should dispense 3 to 8 ml of the liquid to each group.

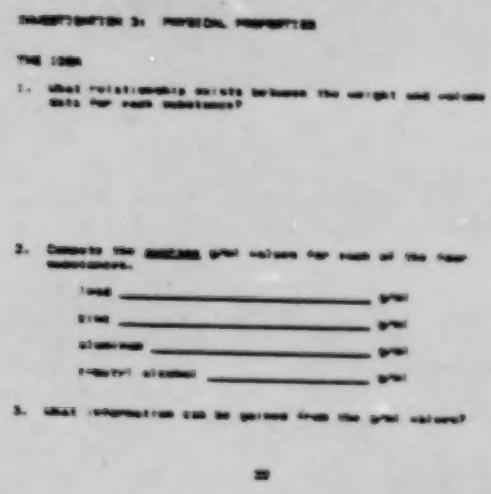
7. The test tube of tert-butyl alcohol stored in Step 7 will be used in Investigation 4.

8. After the students have begun working, make table headings on the board for the weight, volume, and g/ml data for each substance.

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES

CONCEPTUAL INVENTION (THE IDEA)



1. As the weight of the substance increases, the volume increases.

2. Average values should be near the actual values for these substances. Accepted values (taken from density tables) for the densities of these substances are as follows: lead=11.4 g/ml; zinc=7.1 g/ml; aluminum=2.7 g/ml; t-butyl alcohol=0.79 g/ml.

3. The g/ml value tells you the weight of each milliliter of the substance; that is, how much substance is in each milliliter.

INVENTION DISCUSSION

A. In order for the students to graph the weight and volume data for each substance, discuss setting up coordinate axes, labeling, and constant spacing between divisions on each axis. Give examples to show how points are plotted. Discuss the <u>physical significance</u> of the (0,0) point; that is, a <u>sample</u> of zero weight will occupy zero volume.

B. Instruct the students to plot the g and ml data from Step 9 of the Exploration phase.

C. Have the students draw <u>one straight line</u> through each set of data points. Establish that the point (0,0) is common to all sets of data by questioning the students. Each line should be labeled.

BEST COPY AVAILABLE

D. Have the students plot the average g/ml values on the respective graphs.

E. Have the class state how closely these average points came to the lines they drew for each substance. Any lines that are far away from the point can be discussed to determine where the problem lies. In all probability, the error will be the result of having drawn a line that fits poorly through the data points. Ask if the g/ml value could be used to identify a substance.

F. It would be good at this point to have calculated a couple of values for weight and volume (for one or two of the solids used) that would fall on (or near) one of their plotted lines, if the line were extended. Ask the class if a solid sample having those mass and volume data could be identified from their graphs. This activity will help reinforce the 'characteristic physical property' idea.

G. A logical question to ask now is whether weight per unit volume is a characteristic property of gases also.

1. Ask how the values (g/ml) for gases would compare to those for liquids. Have the students explain their answers.

2. In the course of explaining the comparative g/ml values for the three states of matter, students may have used the term "density". If so, ask them to give a functional definition for density. If no one used the word, invent it and lead the class to achieving a clear, concise definition (e.g., the amount of matter per unit volume of a substance).

H. Accepted values (taken from density tables) for the densities of these substances are as follows: lead=11.4 g/ml; zinc=7.1 g/ml; aluminum=2.7 g/ml; t-butyl alcohol=0.79 g/ml.

(a) How do your average values compare to these table values?

(b) Offer a logical explanation for any differences between the experimental values and the <u>table</u> values.

INVESTIGATION 3: PHYSICAL PROPERTIES

EXPANSION OF THE IDEA

[illegible]

Name_____

Section_____

[illegible]

[illegible]

1. [illegible]

2.

	Substance A	Substance B
[illegible]	_____	_____
[illegible]	_____	_____
[illegible]	_____	_____
[illegible]	_____	_____

3. [illegible]

3a

[illegible]

[illegible]

1. [illegible]

2. [illegible]

3. [illegible]

The students should have page 5 to complete while the teacher performs the first demonstration. Page 6 is for the students to fill out during the second demonstration. Both of these demonstrations can be completed in a one hour class session.

PROCEDURE:

Have the class give you instructions for finding out if both of these substances are the same. Obviously, it is possible to use other tests (conductivity, etc.), but hopefully they will suggest finding the density of each of the liquids. If not, the teacher should suggest it.

DISCUSSION:

The densities of the two liquids are different, which will prove that the liquids are different substances. Density is a characteristic physical property which can be used to identify substances. To emphasize the point you could test the liquids for conductivity.

BEST COPY AVAILABLE

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES

[illegible]

Name_____

Section_____

[illegible]

[illegible]

1. [illegible]

2. [illegible]

[illegible]

1. [illegible]

2. [illegible]

This demonstration illustrates density, solubility and miscibility.

PROCEDURE:

1. Pour carbon tetrachloride, water and hexane into a 25 ml graduated cylinder.
2. Pour carbon tetrachloride, hexane, and then water in another 25 ml graduated cylinder.

DISCUSSION:

In graduated cylinder #1, there will be 3 layers of liquids, since when added in this order, each pair of liquids is immiscible (or not soluble), with the other liquid.

The liquids are also layered from least dense to most dense. (Hexane d=0.660 g/ml, water d=1.0 g/ml, and carbon tetrachloride d=1.589 g/ml.)

In graduated cylinder #2 there will only be 2 layers of liquids, since hexane and carbon tetrachloride are miscible.

The students may suggest other ways to add the liquids to the graduated cylinder. Have them offer explanations for their observations.

The density of solid $Pb(NO_3)_2$ is 4.53 g/ml. In the first demonstration you found the density of the $Pb(NO_3)_2$ solution to be around 1.1 g/ml and the density of water was 1.0 g/ml. Emphasize that a solution is a homogeneous mixture of 2 or more substances. $Pb(NO_3)_2$ and water have different densities, yet $Pb(NO_3)_2$ is soluble in water.

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES

C. Assign READING 3: HIERO'S HERO, and the questions. Discuss the points from the reading which need emphasizing and Questions 2b and 3.

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

READING 3: QUESTIONS

SHOW ALL WORK. LABEL APPROPRIATE.

1. A solid brick, weighing 10 grams, has the following dimensions:

   Length = 2 cm; width = 1 cm; height = 1 cm.

   a. What is the density of the brick?

   b. If the brick were cut in half, what would be the density of each piece?

2. You are collecting a gas by displacement of water and the gas is slightly soluble in water.

   a. How does this affect the volume of gas you collect?

   b. How does this affect your determination of the density of the gas? Explain.

3. Draft readings on ships tell how much of the ship is under water. For purposes of this question, assume that the total weight of the ship remains the same and that the ship is not touching the ground. At one time a draft reading of 25 feet-7 inches is made, and at another time a draft reading of 27 feet-1 inch is made. Explain the differences in the draft readings. (Hint: The two readings cannot be taken at the same geographic location.)

Acceptable Answers

1a. V=2 cm x 1 cm x 1 cm=2 cm^3

D=m/V=10 g/2 cm^3=5 g/cm^3

the same

b. V(1/2 brick)=1 cm x 1 cm x 1 cm=1 cm^3

$$D = m/V = \frac{5 \text{ grams}}{1 \text{ cm}^3} = 5 \text{ g/cm}^3$$

b. The gas being soluble in water would not affect the density of the gas. Density is a characteristic physical property.

3. Salt water is more dense than fresh water; therefore, the draft reading would be less in salt water. The draft readings might also be affected by the temperature of the water.

BEST COPY AVAILABLE

INVESTIGATION 3: PHYSICAL PROPERTIES

D. Assign: Density Problems.

[illegible]

Name_____

Section_____

[illegible]

DENSITY PROBLEMS

Solve the following problems. Be sure to show all your work and include the correct units.

1. If a piece of clay has the measurements 2 cm x 4 cm x 6 cm and weighs 96 grams, calculate its density.

2. Nine liters of gasoline have a mass of 6120 grams. What is the density of the gasoline?

3. What is the weight in grams of a block of silver measuring 2.5 cm x 8.0 cm x 4.0 cm, given that the density of silver is 10.5 g/cm³?

4. Calcium chloride has a density of 2.15 g/cm³. What would be the volume of 3.37 g of this substance?

5. If 2 liters of milk weighs 2.06 kg, calculate its density.

Acceptable Answers

1. V=2 cm x4 cm x6 cm=48 cm³

   D=m/V D=96 g/48 cm³

   D=2 g/cm³

2. 9 liters=9000 ml

   D=m/V

   D=6120 g/9000 ml=0.68 g/ml

3. V=2.5 cm x8.0 cm x4.0 cm

   =80.0 cm³

   D=m/V m=VD

   = 80 cm³x10.5 g/cm³

   = 840 g

4. D=m/V V=m/D

   V=3.37 g/2.5 g/cm³

   V=1.57 cm³

5. 2 liters=2000 ml

   2.06 kg=2060 g

   D=m/V

   $D = \frac{2060\ g}{2000\ ml} = 1.03\ g/ml$

BEST COPY AVAILABLE

BEST COPY AVAILABLE

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

TEACHER'S GUIDE

INTRODUCTION

In this investigation we will center our attention on the process of changing the phase of a substance. Before beginning this investigation a discussion of the three phases of matter is desirable. The following may be useful in generating an outline for the discussion.

There are three phases (or states) of matter which we commonly observe and with which we commonly interact. Those phases are the solid phase, the liquid phase, and the gas phase. . Under normal environmental conditions, some substances are found in only one of the three phases. For example, gold, calcium carbonate, and sulfur are solid phase substances, but mercury is a liqud and nitrogen and carbon dioxide are in the gaseous phase. Water, by contrast, commonly exists in all three phases.

What determines the phase in which a substance will be found? Considering water, it is apparent that the temperature of the environment is an important determining factor in the phase of a substance. The lower the temperature, the more likely a substance is to be found in the solid phase. Conversely, a higher temperature would make it more likely to find a substance in the liquid or gaseous phase. Furthermore, to change a liquid to a solid, it is necessary to remove heat from that liquid. A gas can be formed from a liquid by adding a sufficient amount of heat to the liquid.

An interesting and very important point about these changes is that they are physical changes. When a substance freezes or melts or boils or condenses there is no change in the chemical nature of the substance. Water is water whether it is in the form of ice (solid), liquid water, or water vapor (gas).

We will be using tert-butyl alcohol in our investigation, although it is important to realize that the same principles apply to almost every pure substance.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

The conceptual inventions are melting point and boiling point. The melting point is defined as the temperature at which a solid changes to a liquid while heat is being added. The temperature at which a liquid changes into a gas while heat is being added is the boiling point. Stress the fact that melting point and boiling point are <u>temperatures.</u>

During the Expansion of the Idea phase of this investigation the following terms will be defined and discussed: temperature, latent heat of fusion, latent heat of vaporization, pure substance, mixture, and intensive and extensive physical properties.

CONCEPT

<u>Melting point</u> is the temperature at which a solid changes to a liquid, or a liquid changes to a solid, while the temperature remains constant. <u>Boiling point</u> is the temperature at which a liquid changes to a gas, or a gas changes to a liquid while the temperature remains constant.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

EQUIPMENT

A. Exploration

1. Labware

a. 10 ml graduated cylinder
b. 18x150 mm test tube
c. 400 ml beaker
d. -10 to 110°C thermometer
e. ring stand with iron ring
f. wire gauze with asbestos center
g. utility clamps
h. laboratory burners
i. boiling chips

2. Reagents

a. tert-butyl alcohol (stored from step 7 of Investigation 3)
b. ice

B. Expansion of the Idea

1. Labware (same as Exploration)

2. Reagents

a. para-dichlorobenzene
b. cyclohexane
c. n-propanol
d. naphthalene
e. methanol

PRE-LAB INSTRUCTIONS

A. Assign the sample size each group is to use in Step 1, Page 1. This will give the class a spread of volume data during the discussion. A 5 ml sample takes a considerably long time to boil away during the boiling point determination. The larger sample size will, however, give a longer plateau on the melting point determination.

B. Instruct the students to stand <u>upwind</u> of the systems during the heating (boiling procedure).

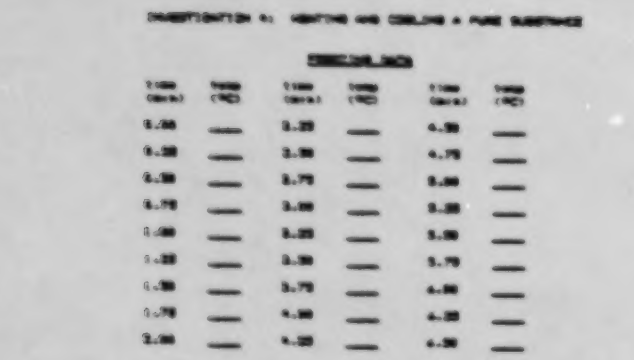
C. When heating, the laboratory burners should be left underneath the beaker throughout the procedure.

D. Make sure that the students understand what is expected of them during these procedures. As necessary, demonstrate various parts of the procedures using a mock-up of the experimental apparatus.

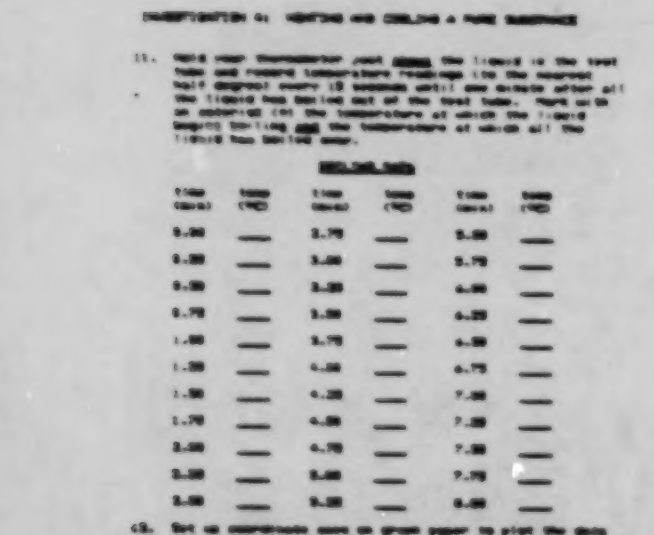
BEST COPY AVAILABLE

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

EXPLORATION



If necessary, students can extend the time/temperature data table to contain all their data.



Guidelines for graphing the data can be given to the students after they have taken data. A smooth curve should be drawn through the data points. It may be impossible to draw the curve through all the data points.

Discuss the lack of physical significance of 84°C temperature on the graphs.

BEST COPY AVAILABLE

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

CONCEPTUAL INVENTION (THE IDEA)

[redacted]

Acceptable Answers

a. liquid b. liquid-solid c. solid

a. liquid b. liquid-gas c. gas

Additional information for tert-butyl alcohol (2-methyl 2-propanol)

m.p.=25.8°C

b.p.=82.41°C

3. The freezing temperature is the same regardless of the sample size.

4. The boiling temperature is the same regardless of the sample size.

INVENTION DISCUSSION

A. Heating Graph Discussion

1. Have the students describe the general shape of the curve.

2. Center their attention on the heat-up section of the curve.

   a. Ask why the temperature is increasing during this time. (Heat, or energy, is being added to the system.)

   b. Ask what the appearance of the liquid was during this time (wavy movements in the liquid).

BEST COPY AVAILABLE

3. Now center the questioning on the flat portion of the graph.

   a. Ask what the appearance of the liquid was during this time. (Liquid was boiling.)

   b. Ask if the system was being heated at the same rate as before. (Yes, the burner was still under the water bath.)

   c. Ask why the temperature did not change during this time period (energy being used in the change).

NOTE: This is a deep question and the teacher need not expect an immediate response. The teacher should lead the discussion carefully, asking questions whose answers will lead the class to the idea that the energy is showing up as a phase change instead of a temperature increase. This discussion can be taken to the point of mentioning that forces of attraction between particles of the substance (molecules) must be overcome. Furthermore, some discussion of freedom of molecular motion (liquid compared to gas) is appropriate at this time. Such discussion can be enhanced during the later discussion of melting point.

4. Invent <u>boiling point.</u> The boiling point is the temperature at which a liquid changes into a gas while heat is being added.

B. Cooling Graph Interpretation

1. The discussion of the cooling data should follow the same guidelines as the heating data discussion.

2. Invent <u>melting point</u> as the term that identifies the temperature at which the solid to liquid (or vice versa) phase change occurs. Although <u>freezing point</u> would seem to be equally adequate and even more accurate in a cooling operation, the term <u>melting point</u> is most often used in chemical literature.

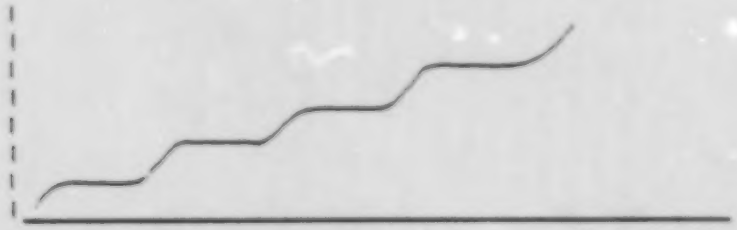
3. The location of the asterisks, with reference to the flat portions of the graphs, is a matter that warrants some discussion.

   a. Ideally, these asterisks would appear at coordinates corresponding to the beginning and the end of the flat portion: that is, the portion of the procedure during which both phases existed.

b. Have the class supply reasons for the deviations from ideality. The students will probably provide suggestions that can be molded into a coherent explanation. If the suggestions are summarized on the board as they are supplied, the teacher will be better able to help the students formulate such a coherent explanation. (The essence of these responses should relate to uncertainty in exact time observations for these changes, or the deviations may be due to lack of ideality in the manipulation of the systems, such as positioning of the thermometer or inadequately vigorous stirring.)

C. Summary Discussion

1. Sketch a heating curve on the board that has <u>four</u> plateaus. Have the students dictate what each portion of the curve represents; for example, solid being heated, solid and liquid, etc. The top and bottom plateaus could refer to the gas and solid phases (respectively) being at the same temperatures as their environments.



Ask your students the following questions concerning the unlabeled graph.

a. What is happening on this graph? (The substance is heated.)

b. Why does the temperature stay constant in flat portion #3? (Change of phase or temperature equilibrium with the environment.)

c. When does an increase in temperature occur? (Only when one phase is present.)

d. What phase(s) is/are present in each portion?

e. Invent 'latent heat of fusion. Latent heat of fusion is the energy removed to change a liquid to a solid at the melting point (or vice versa).

f. Invent latent heat of vaporization, which is the energy added to change a liquid to a gas at the boiling point (or vice verse).

2. It is strongly recommended that the teacher perform a <u>freezing</u> procedure for paraffin wax and plot the data on an overlay. Present this graph to the class and ask for an explanation of the graph. (There is no plateau, therefore there must be no melting point.)

NOTE: Paraffin wax is suggested because of its complete lack of a freezing plateau.

a. Point out that paraffin is a mixture of a number of substances, so that it does not behave the way a single pure substance would.

b. Be sure that it is clear to the students that if a m.p. or a b.p. is given for a substance, those data are referring to the characteristics of the <u>pure</u> substance.

c. Show the class that the elements on the periodic table have melting points and boiling points listed for them. You might want to show the class which elements are gases, which are liquids, and which are solids at room temperature.

3. Ask the class how the density, m.p., and b.p. of pure substances vary with amount of substance. When the class has concluded that amount has nothing to do with the values for these characteristic properties, invent the term <u>intensive physical property</u> to describe these physical properties.

a. There are several other intensive physical properties that can (and should) be elicited from the class (preferred) or given by the teacher. These include: heat and electrical conductivity, malleability, ductility, etc.

b. Any physical properties that depend on the amount of substance are terms as <u>extensive physical properties</u>". (Examples include length, weight, volume, area.)

EXPANSION OF THE IDEA

A. Experiment

[illegible]

Name_____

Section_____

[illegible]

[illegible]

1. [illegible]

[illegible]	[illegible]	[illegible]	[illegible]
[illegible]			
[illegible]			____
[illegible]			____
[illegible]			____
[illegible]			____
[illegible]			____
[illegible]			____
[illegible]			____
[illegible]			____

[illegible]

2. [illegible]

[illegible]					

1. Discuss, with student input, the procedures that should be used to determine the unknown.

   Inform the students that they do not need to boil the entire sample, but use temperature stabilization as the boiling point. Get students to verbalize this idea by questioning; for example, "How would you know when you have reached the m.p. or b.p. of your substance?"

   The only substances from the list that are acceptable for testing are para-dichlorobenzene, cyclohexane, n-propanol, naphthalene, and methanol.

BEST COPY AVAILABLE

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

B. Assign READING 4: WIGGLES, VIBRATIONS, AND THE TEMPERATURE OF THINGS, and its questions. The reading discusses the kinetic molecular theory on an elementary level. Discuss the reading with the students stressing the significant aspects. Invent <u>temperature.</u>

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

Name _____ A

Section _____

READING 4: QUESTIONS

1. If a glass of water is left uncovered at room temperature, it will eventually all evaporate. Given that the boiling point of water is 100°C, offer an explanation for the evaporation of the water.

2. What characteristic of a substance will determine what the value of its boiling point will be? Give information that will logically support your answer.

3. Explain why dew is found on grass in the early morning instead of in the early evening.

4. Explain how it is reasonable to assume that the particles within solid objects are in motion.

1. Some molecules at the surface have enough energy to break away and change phase from a liquid to a gas. This is possible because temperature is a measure of the <u>average</u> kinetic energy of a substance. Some particles, therefore, will have sufficient energy to escape the liquid.

2. *This is a difficult question. If this is a sheet to be graded, consider not grading this question.

   (When a substance is boiled, energy is needed to break the molecules apart. Different substances require different amounts of energy to separate the molecules because of the unique attractions among their molecules.

3. The ground cools by radiating heat to the atmosphere. When the temperature of the ground reaches the <u>dew point</u> for the existing water vapor concentration, water will precipitate from the air that comes in contact with the ground.

4. Temperature is a measure of the average kinetic energy (motion energy) of the particles in a system. Since a solid object has a measurable temperture, we can assume that the particles of that solid are moving.

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

TEACHER'S GUIDE

INTRODUCTION

This investigation has as its major objective the development of the need for an atomic model for matter. This objective will become apparent to most students only in retrospect; that is, after the atomic model has been developed (see Investigation 6). The results of the major work done in this investigation--the determination of a compound formula through experimentation--serves a vital purpose. It leads to a preliminary exploration of the periodic table to ascertain values (with no units for the time being) for the weights of elements. Reaction terminology, reactants and products, will be used and synthesis reaction is invented during the discussion.

The Expansion of the Idea laboratory (student sheets 5 through 8) has several possible uses. As the learning cycle is written, this laboratory is another opportunity for the students to determine experimentally the formula of a compound. The materials used in this laboratory are less expensive than those required in the Exploration phase. If the cost of materials is a factor, the Expansion laboratory could be substituted as the Exploration.

CONCEPT

Compounds have definite proportions of weight and numbers of atoms.

EQUIPMENT

A. Exploration

1. Labware

a. laboratory balances
b. evaporating dishes and crucible covers
c. ring stands with small iron rings
d. pipestem triangles
e. laboratory burners
f. crucible tongs
g. fine steel wool
h. scissors

BEST COPY AVAILABLE

2. Reagents

a. magnesium metal ribbon
b. distilled water

B. Expansion of the Idea

1. Labware

a. 18 x 150 mm test tube
b. laboratory burner
c. test tube clamps
d. evaporating dish
e. laboratory balance

2. Reagents

a. hydrated copper (II) sulfate
b. distilled water

PRE-LAB INSTRUCTIONS

A. Remind the class about the techniques of gentle and strong heating. This is especially important during the second, third, or fourth heating of the system because of the potential of breaking the evaporating dish.

B. The teacher should provide magnesium ribbon cut in strips 40 to 50 cm long. This length range is recommended because it allows an easily measured weight gain at the end of the reaction. If different groups get different lengths of ribbon, a desirable range of weights will also result.

C. Instruct the students to cut the cleaned ribbon into small pieces (approximately 1 cm). This configuration will allow for more uniform heating of the metal promoting more complete reaction during the initial heating.

D. There should be no reason for lifting the crucible cover off the system during heating. The metal should be placed entirely under the crucible cover. Instruct the students to stop heating the system temporarily if white smoke starts coming out of the system. The reason for this caution is to prevent weight loss from the system.

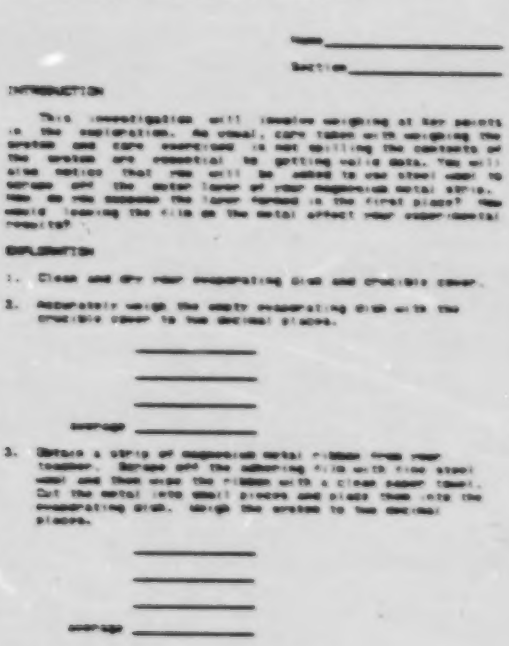
BEST COPY AVAILABLE

E. After the water has been added, caution the students to heat the systems very gently until all the water has boiled away. The crucible cover can be held just above the mixture, using crucible tongs to prevent material from spattering out of the system during the heating process.

F. The teacher will need to explain the idea of heating to constant weight. The goal of this weighing is to obtain <u>two consecutive average weights</u> that are the same, within the uncertainty of the balance. The students should be instructed to repeat the procedure in Step 7 until the constant weight is obtained.

G. The systems must always be cool before weights are taken. This is to prevent damage to the balances.

EXPLORATION



INVESTIGATION 5: COMPOSITION OF A COMPOUND

Name_____

Section_____

INTRODUCTION

This investigation will involve weighing at key points in the exploration. As usual, care taken with weighing the system and care exercised in not spilling the contents of the system are essential to getting valid data. You will also notice that you will be asked to use steel wool to scrape off the outer layer of your magnesium metal strip. How do you suppose the layer formed in the first place? How would leaving the film on the metal affect your experimental results?

EXPLORATION

1. Clean and dry your evaporating dish and crucible cover.

2. Accurately weigh the empty evaporating dish with the crucible cover to two decimal places.

_____

_____

_____

average _____

3. Obtain a strip of magnesium metal ribbon from your teacher. Scrape off the adhering film with fine steel wool and then wipe the ribbon with a clean paper towel. Cut the metal into small pieces and place them into the evaporating dish. Weigh the system to two decimal places.

_____

_____

_____

average _____

3. The metal pieces should be approximately 1 cm in length.

5. Remind the students not to remove the lid during the heating process.

Ask the students to suggest reasons that heating should be stopped if white smoke begins coming from the system.

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

4. Using your data, calculate the weight of magnesium metal in the crucible. Show your work.

5. Heat the covered evaporating dish and magnesium crucible gently for two minutes, gradually increasing the heat intensity. Continue heating the crucible strongly for an additional 10 minutes, then turn off the burner.

6. As the crucible is cooling, critically inspect the contents of the crucible. Record your observations.

7. After the evaporating dish has cooled, add very enough distilled water to the dish to cover its contents. Gently heat the evaporating dish. As the water boils away, watch the vapors coming from the evaporating dish and identify them. After the water has evaporated, continue to strongly heat the crucible for five minutes. Heat to constant weight.

_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

8. Determine the weight change in the crucible. Show your work.

9. Record careful observations of the final contents of the evaporating dish.

10. You have now calculated two weights: the weight of magnesium and the weight gained (or lost) during reaction. There are four mathematical operations (addition, subtraction, division, multiplication) available to manipulate these weights. Use these operations on the weights and fill in the chart below. Include units, when appropriate.

g Mg + g O	g Mg − g O	g Mg / g O	g Mg × g O

CONCEPTUAL INVENTION (THE IDEA)

THE IDEA

1. What substances could possibly have reacted with the magnesium? (That is, what are the substances available to the crucible?

2. Which of the four mathematical operations in Step 10 would you expect to give the same value, no matter how much Mg was used? Why?

Acceptable Answers

1. O_2, N_2, CO_2, $H_2O(v)$, H_2, Ar, He, and any other gases that they desire to list.

2. Accept the mathematical operation they suggest if a reasonable explanation is given.

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

INVENTION DISCUSSION

One of the central ideas of this investigation is that element weights are unique and useful in determining compound formulas. To this end the students have been instructed to manipulate the data according to the four basic mathematical operations. The discussion of this investigation begins, however, by examining--from experimental evidence and logical reasoning--the possible results of the reaction that took place. Guidelines for this discussion follow.

A. Ask the students if a chemical change occurred in the crucible (during the first heating). List the evidence for a chemical change. The observations listed here should relate to what happened before water was added.

B. If magnesium underwent a chemical change, what reacted with the magnesium? Ask the class to supply a list of reactants available to the magnesium. Write these on the board in equation form as they are suggested (for example, $Mg+O_2$-->). It can be noted that carbon dioxide and water vapor are abundantly available from the burner flame.

C. When the list is completed, lead the class in predicting products of these reactions and then in balancing the equations. Formulas for some predicted products will need to be supplied by the teacher while others, such as that of magnesium oxide, can be given by students after they have looked them up or recalled them from previous investigations.

NOTE: Initially, all possible reactants should be given consideration. Some, such as argon, can soon be dismissed due to their "noble" nature, but others, such as carbon dioxide, should be given a more detailed treatment. Allow the class to predict products (in terms of which elements are combined in compounds) and to support their attempts with some sort of evidence. For example, some students have suggested that since there was some black solid present after the heating, carbon dioxide could have reacted with the magnesium to form magnesium oxide and carbon.

D. The net result of the foregoing discussion should leave only equations producing magnesium oxide and magnesium nitride, as follows:

$$2Mg+O_2 \rightarrow 2MgO$$

$$3Mg+N_2 \rightarrow Mg_3N_2$$

E. Ask the students what they know about MgO--see Investigation 1 (white powder). Did all of the magnesium ribbon react? (Probably not.)

F. Ask the class the purpose for adding water to the reacted system. If the students have trouble with that question, ask for their observations as they heated the system to drive off the water. Lead the discussion to an equation representing the reaction of magnesium nitride and water to produce ammonia and magnesium oxide. By this time in the discussions of the reactions, it should be apparent to the students that all reaction possibilities eventually lead to the formation of magnesium oxide.

G. Using the equation that reacts magnesium with oxygen to produce magnesium oxide, invent <u>synthesis reaction.</u>

H. Make column headings on the board that correspond to the ones in step 10 of the Exploration. Ask a member from each group to record the information from that group on the board. When all the data are listed and copied down by the students, lead a discussion about the information provided by the four data calculation. A discussion of units is appropriate.

1. What does the sum weight represent? (The weight of the product.)

2. What is the product? (MgO.)

3. What does the weight change represent? (The weight of oxygen.) From this point on, the weight change can be referred to as oxygen weight.

4. What information does the difference provide? (The magnesium in MgO weighs more than oxygen.)

5. Question the students about the meaning of the ratio (comparison). Be sure that the students have some understanding of this relationship before continuing.

6. What information does the product provide?

   (Nothing.) What is the physical significance of g^2? (None.)

I. The students have had to accept the formula for magnesium oxide on the teacher's authority. Ask the class if these data ratios can be used to show that MgO is a proper formula for magnesium oxide. Through discussion and questioning, present the idea that the periodic table lists weights for one/unit/of each element. Ask the class to calculate the weight ratio of one magnesium and one oxygen from the weights given on the periodic table.

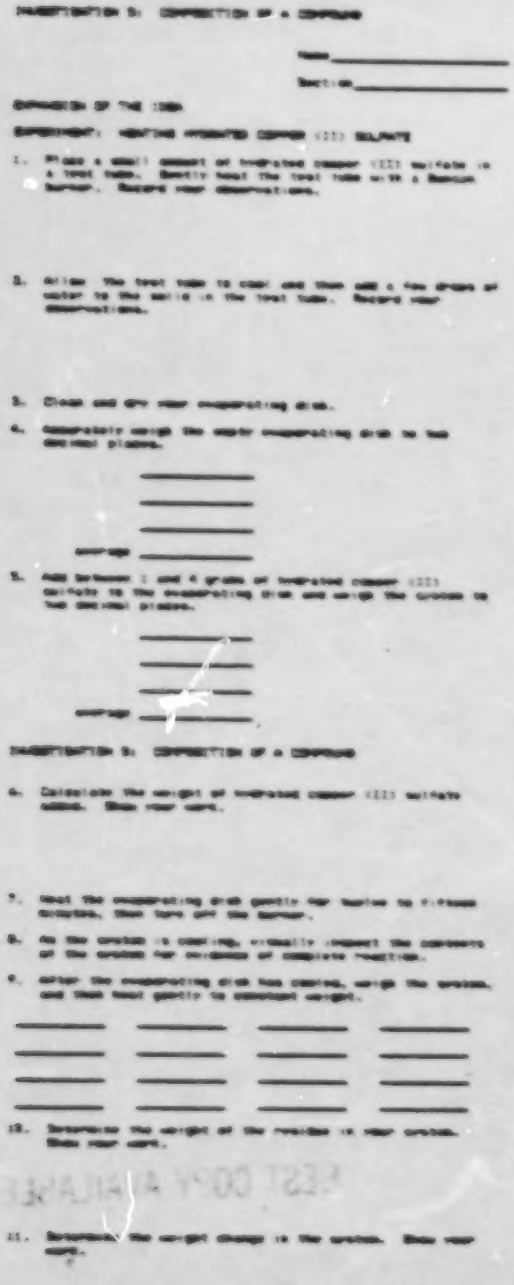
J. Compare the table ratio obtained above to the experimental ratios. Discuss discrepancies, average the experimental ratios, etc., but conclude the discussion by gaining consensus on the viability of the formula.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

EXPANSION OF THE IDEA

EXPERIMENT

A. Review sugestions in the Introduction concerning the use of the laboratory in this section of the learning cycle.



1. Give students powdered samples. The students should use a test tube clamp to hold the test tube horizontally.

5. A spread of weights is needed to make the class data easier to interpret. The teacher may want to insure the spread by assigning weight ranges to different groups. The hydrated copper (II) sulfate crystal should be powdered.

7. There should be no blue color remaining in the sample. If there is, heating should be continued until no blue color remains.

9. The students should be reminded to heat gently and be reminded what is meant by <u>heating to constant weight.</u>

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

[illegible]

[illegible]

1. The chemical formula for hydrated copper (II) sulfate can be written

   $CuSO_4(H_2O)_x$,

   where x represents a number of H_2O molecules. What are the possible products of heating this substance? Write equations for the various reactions predicting.

2. Did the weight increase or decrease when you heated the hydrated copper (II) sulfate:

   $CuSO_4(H_2O)_x$?

   Offer an explanation for the change.

3. Divide the weight change by the weight of the residue. Show your work and explain what this number represents.

4. Use the weights on the periodic table to determine the best value of "x" in

   $CuSO_4(H_2O)_x$.

   x should be an integer. Show your work.

A class discussion of the questions will be necessary.

1. Have the students list the possible products of heating $CuSO_4(H_2O)_x$. The students should use laboratory evidence to support the possible products.

2. The weight change represents the weight of water. Again, the students should use laboratory evidence to explain the change in weight.

3. Discuss the significance of the constant value for the weight change divided by the weight of the residue. Stress the definite proportion concept for both weight and number of particles.

4. Work through the problem and discuss its significance in terms of specific formulas for chemical substances.

Discuss the laboratory. Possible questions follow.

1. List observations made while heating copper (II) sulfate in a test tube.

2. What happened when water was added to the test tube?

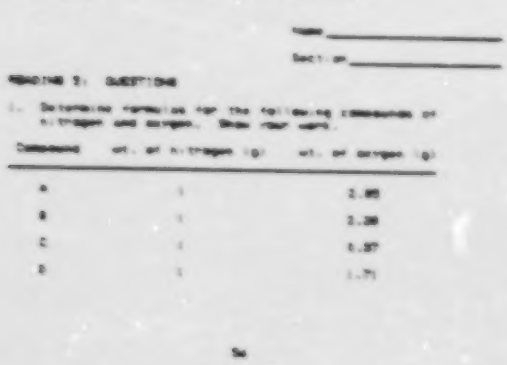
3. Describe the residue in the evaporating dish after heating.

4. The teacher may want to challenge the students' assertion that the liquid which appears in the test tube is water. $CoCl_2$ paper be introduced during the discussion.

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

B. Assign Reading 5: THE CHEMICAL CONGRESSMAN, and the questions at this time. You may wish to lead the class through the formula determination for one of the compounds in Question 1. The questions and acceptable answers are as follows.



1. Formula will be N_xO_y. Use u for units.

Compound A

$$\text{Assume } N_2O_5 = \frac{5(16u)}{2(14u)} = \frac{80}{28} = 2.85$$

Compound B

$$\text{Assume } NO_2 = \frac{2(16u)}{1(14u)} = \frac{32}{14} = 2.28$$

Compound C

$$\text{Assume } N_2O = \frac{1(16u)}{2(14u)} = \frac{16}{28} = 0.57$$

Compound D

$$\text{Assume } N_2O_3 = \frac{3(16u)}{2(14u)} = \frac{48}{28} = 1.71$$

2. Source A: CO_2

$$\frac{13.3\ g}{5.0\ g} = 2.66$$

$$\frac{2(16u)}{1(12u)} = 2.66$$

Source B: CO_2

$$\frac{8.8\ g}{3.3\ g} = 2.66$$

$$\frac{2(16u)}{1(12u)} = 2.66$$

BEST COPY AVAILABLE

INVESTIGATION 5: COMPOSITION OF A COMPOUND

INVESTIGATION 5: COMPOSITION OF A COMPOUND

4. Which of the three laws discussed in this reading are illustrated by the heating of magnesium in Investigation 5? Explain.

5. Which of the three laws discussed in this reading are illustrated by the heating of hydrated copper (II) sulfate in Investigation 5? Explain.

3. Question 1 illustrates the Law of Multiple proportions.

   Question 2 illustrates that the same substance can be present in different situations--Law of Constant Composition.

4. The Law of Constant Composition is illustrated by the heating of magnesium. The ratio of the weights involved were a constant value.

5. The Law of Conservation of Mass was illustrated by the heating of hydrated copper (II) sulfate. The weight of the water loss plus the weight of the residue equaled the weight of the hydrated copper (II) sulfate.

BEST COPY AVAILABLE

INVESTIGATION 6: ATOMIC THEORY AND MATTER

TEACHER'S GUIDE

GENERAL INTRODUCTION

Since this series of discussions will constitute the first formal presentation of the atomic model concept in this course, the applicability of the model to chemical explanations will be limited. It will, however, allow for simple, logical explanations of A-column ion formation and specific formula representations for many compounds. Once the model is understood and applied to binary compounds of A-column elements, the generalization to more complex compounds--those whose formulas can be obtained by using common polyatomic ions--can be made quite easily.

The proper placement of this investigation could be a matter of considerable debate. One could argue that this series of discussions, excluding the preliminary written exercises, could be presented at the beginning of the course with no prior learning required. It is the opinion of the authors, however, that such an approach would essentially ignore the readiness of the students to deal effectively with the ideas involved. This investigation, therefore, has been placed so that the students will have had time to interact with several chemical systems that are designed to prepare them for a fruitful study of the atomic model.

The intent of the WRITTEN EXERCISES: ASSUMPTIONS is to encourage the students to make a careful analysis of the work that they have done up to this point in their study of chemistry. Many concepts that will be invented will have their understanding enhanced by relating them to the work that has been done earlier in the course. The teacher should request a minimum of twenty assumptions from each person (one student wrote 84 assumptions). Students have responded well to the encouragement to begin at Investigation 1 and go through each investigation, carefully making a list of assumptions made in each one.

The format of this investigation is quite different from that of the preceding ones. The reason for this change is that there is essentially no laboratory work within this investigation--the major activity is class discussion. The teacher would be wise to place strong emphasis on student contributions in the course of these discussions. More detailed suggestions for incorporating student inputs are included in the applicable discussion guidelines.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

In the content summary which follows, the concepts that are to be invented will be listed under the discussions in which they are to be developed. It is important to note that further refinements of the atomic model will be made as they are needed (for example, electron energy sublevels and orbitals--see Investigation 18). It is, therefore, recommended that the teacher be patient with the development of the model.

CONTENT SUMMARY

NOTE: These discussion guides are divided according to topic coverage and not according to time blocks. The teacher should not be concerned if a given discussion requires two or three class periods to complete.

A. FIRST DISCUSSION: WHAT IS AN ATOM?

1. Nucleus
2. Proton
3. Neutron
4. Electron
5. Atom definition
6. Isotope (optional)
7. Atomic number (optional)

NOTE: These last two ideas have been listed as optional because their invention in the <u>first discussion</u> is not essential. In some cases, discussion would have to be forced to allow for the invention of one or both of these ideas at this time. Other situations, however, lend themselves very well to these inventions. At any rate, both concepts <u>can</u> be saved until the fourth discussion.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

B. SECOND DISCUSSION: HOW DO THE PARTS OF AN ATOM INTERACT?

1. Types of electrical charge
2. Interaction of like charges and unlike charges
3. Electric force field

C. THIRD DISCUSSION: APPLICATIONS OF CHARGE INTERACTIONS TO THE ATOMIC MODEL.

1. Description of nuclear forces
   a. Short-range
   b. Charge independent
   c. Strong
   d. Attractive
2. Total electron energy (potential plus kinetic)
3. Electrical neutrality of atoms
4. Interatomic electrical effects

D. FOURTH DISCUSSION: ELECTRON ENERGY.

1. Positive and negative ion formation
2. Atomic number (if not previously invented)
3. Relative electron energies
4. First ionization energies
5. Isotopes (if not previously invented)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

E. FIFTH DISCUSSION: PATTERNS IN FIRST IONIZATION ENERGIES.

1. Period of elements
2. Family of elements
3. Periodic property
4. Second ionization energy

F. SIXTH DISCUSSION: PATTERNS IN SECOND IONIZATION ENERGIES.

1. Electron energy levels

G. SEVENTH DISCUSSION: INTRODUCTION TO CHEMICAL NOMENCLATURE.

1. Valence level
2. Noble gas electron configuration
3. Charge determination of specific ions
4. Ionic bonding
5. Formulas and compound naming

INVESTIGATION 6: ATOMIC THEORY AND MATTER

EQUIPMENT

1. Electroscope
2. Glass and rubber excitation rods
3. Metal and glass rods for heating experiment (optional)
4. Silk cloth
5. Two pith balls on strings
6. Two rings with ringstands

I. WRITTEN EXERCISES: ASSUMPTIONS

A. This activity can be begun in one class period with the remainder being assigned as homework. By having the students start working on it in class, the teacher will be able to give the students additional help in making their lists. If the teacher has the students list the assumptions by investigation, the students will be able to organize their lists better. Furthermore, this procedure will facilitate the teacher's examination of the list.

B. The purpose of this exericse is two-fold:

1. Students are exposed to the importance of making reasonable assumptions; and
2. Students involve themselves in a careful review of the work they have done to this point in the course.

C. Since the content review will be important in the development of concepts in the present investigation, it is recommended that the assumptions exercises be graded.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

II. FIRST DISCUSSION: WHAT IS AN ATOM?

NOTE: The following method for building a model for the atom utilizes what students have heard or read about atoms previously, instead of an initial historical approach to the model. The main reason for this is that we have found that the historical approach is of little interest to the majority of students at this point in their study. Therefore, the historical approach has, at best, a nominal pedagogical value initially. After the model has been constructed and reasonably refined by the following method, the historical aspects of the development of atomic theory should have more meaning, and the teacher may be able to encourage the students effectively to delve into these aspects.

A. Pose this question (What is an atom?) to the class in such a way that they will give you information from their past experience (science classes, press, etc.).

1. List all information on the board. It is from this information that we will begin constructing our model of the atom, so the teacher can accept anything (within reason) that they say. Depending on how much terminology they know, we may not be able to use everything that is listed when constructing the initial model, but the teacher can provide the direction for the use of the information to develop the model in an orderly manner.

NOTE: It may be that many students are initially reluctant to offer information for the list. Experience has shown that this reluctance diminishes sharply when the students realize that the teacher is willing to accept their contributions. Encourage the students to write down the list of information. Also, you may want to have one student either loan you his/her list (so that you can reproduce a copy to keep) or write out the list for you. The teacher should use this information, when it is applicable, throughout the development of the model. Whenever possible, this information should be used in the form in which it was contributed by the class. This procedure allows the students to be involved as contributors to the development of the model.

2. The first model to be represented on the board, as a synthesis of the information given, should be one that includes a <u>small,</u> positively charged nucleus that is surrounded by a <u>relatively large</u> negatively charged cloud. Obviously, the information contributed will have to include the essence of this model. It is the teacher's responsibility to ensure that the presentation (invention) of this diagram is properly timed. Example:

3. We need to make it clear--from the beginning of our study--that the atom occupies <u>three dimensional space,</u> even though most representations on paper can only show two dimensions. This point should be made frequently.

4. It is reasonable to ask what good this representation can do for us, with reference to

   (a) what an atom is;

   (b) how its parts interact with each other; and

   (c) how different types of atoms interact with each other.

   These three reference points are important to list on the board because they will be the basis for refinements to the model. The listing will also provide a direction for the students' thought during subsequent discussions. These items can be brought in logically as the discussions continue.

5. The present model will permit the atom to be defined as <u>the smallest particle of matter that can be distinguished as an element.</u> In arriving at a definition of this sort, the students should make the maximum possible contribution. The teacher's responsibility is to synthesize the student input into a clear, coherent definition.

NOTE: It is good to write all definitions on the board. Students should be encouraged to copy these definitions into their notebooks.

6. The establishment of a definition for the atom completes this initial discussion. To determine what this representation shows about the interaction of the parts of an atom, we will have to look at some characteristics of these parts (that is, the charges) that we say make up the atom.

III. SECOND DISCUSSION: HOW DO THE PARTS OF AN ATOM INTERACT?

A. This discussion will center on the way that static charges behave. To accomplish the discussion, a series of demonstrations, using an electroscope, will be performed by the teacher. The teacher must always be alert to the necessity of maintaining continuity during this discussion and to the necessity of relating this discussion to the previous discussion. The demonstrations should be practiced by the teacher before presenting them to the class.

B. Before the demonstrations can have much meaning for the students, the students need to have a firm idea about which materials will allow passage of electrical charge and which materials will not allow the passage of electrical charge. It is because of this that the following <u>optional</u> activity is included.

1. Show the class a metal rod and a glass rod having the same dimensions.

2. Ask which rod would heat the fastest, if they were both placed in a burner flame.

3. Since there is usually disagreement about which one will heat the fastest, the teacher can suggest that the class devise an experiment to show which one will heat more rapidly. It is good to write the procedure so devised on the board.

4. When the students are satisfied that the procedure is valid and properly controlled, ask for four volunteers--two to heat glass rods and two to heat metal rods. On signal, these students will heat the rods for a predetermined amount of time.

5. At the end of this heating period, ask the experimenters to determine which rods got the hottest. It will be readily determined that the metal rods got hotter over a greater length than did the glass rods.

6. Ask the class how well the different rods would conduct electricity. Most students will be sure that the glass will not conduct and that the metal will conduct. The teacher can show this to be so by setting the rods on a non-conducting table and touching first one and then the other with both electrodes of a light bulb conductivity apparatus. (Caution: Do not hold the metal rod while testing its conductivity.)

7. It remains for the teacher to make the generalization (with help from the class) that metals are generally good conductors of heat and electricity and that non-metallic substances are not good conductors of heat and electricity.

8. Point out the parts of the electroscope with reference to the foregoing experimentation and discussion.

C. The teacher needs to invent the idea that there are two kinds of electrical charge--positive and negative--and that the charges in the atom are indeed electrical charges. These points are often brought out during the first discussion, and if that occurs, it would still be wise to mention them again at this time. It is, in fact, a good opportunity to have the students refer to the lists that they have made during the first discussion.

1. After stating that an electroscope is an instrument used to detect electrical charge, bring an uncharged glass rod near the electroscope to show that it has no effect on the leaves.

a. Ask the class to explain why the leaves did not move. (The rod is neutral; that is, it has the same amount of positive charge as negative charge.)

2. Induce a charge on the glass rod by rubbing it briskly with a piece of silk. Bring the charged rod near the electroscope and note the effect of the charged rod on the leaves of the electroscope. Move the rod away from the electroscope to allow the leaves to collapse.

a. Ask the class to tell where the charge is (on the glass rod). When this information has been obtained, ask if any charge was transferred from the rod into the electroscope. (No, the leaves collape as soon as the rod is taken away.)

3. If the glass rod still contains a sufficient amount of charge, touch the glass rod to the electroscope so that enough charge will be transferred to the electroscope to keep the leaves spread well apart. (You may need to rub the rod again to get enough charge to do this properly.)

a. Develop, through questioning and discussion, the idea that the effect of a charge extends well beyond the matter that contains the charge. This can be done by referring to the two demonstrations you have just performed. (The idea has obvious value, when it is later related to the effective space occupied by electrons, and to the effect that the nucleus of one atom can have on the electrons of another atom.) Invent electric force field, and define it as the space around a charge that is affected by the charge.

b. If you have not already told the class that the charge on the rod is positive, do so, and then ask if the same charge is now trapped on the electroscope (yes).

4. Move the glass rod near the electroscope again.

a. Ask what this demonstration shows about the way positive charges interact (positive repels positive).

b. Ask how distance affects the force due to the charge (the effect decreases as the distance increases).

5. You are now ready to introduce a negative charge to this system. The negative charge can be produced by rubbing a rubber rod with wool or rubbing it on your hair if it is dry. After telling the class that you have produced a negative charge on the rubber rod, bring it <u>near</u> the electroscope.

   a. Question the class about what they have seen to develop the idea that the demonstration shows that unlike charges attract each other.

6. Discharge the electroscope by touching the metal ball with your finger.

   a. Have the class predict the way negative charges would interact with the discharged electroscope and how negative charges would interact with other negative charges.

   b. After the predictions are submitted, perform the procedures done earlier with the glass rod, but use the rubber rod instead.

   c. Summarize the material covered in this demonstration-discussion on the board so that the class will be encouraged to record the information in their notes.

      (1) positive charge repels positive charge

      (2) negative charge repels negative charge

      (3) positive and negative charges attract each other

      (4) the effect of a charge is noticed in a space or field around the charge--in the electric force field.

D. Relate this information on charges to the model of the atom that has been developed. Discussion of this point should be guided so that the students understand that the positive and negative static charges were developed by the transfer of electrons only. This information raises several questions about the model that should be posed to the class.

1. Since like charges repel, why doesn't the nucleus fly apart?

2. Since unlike charges attract, why aren't the electrons pulled into the nucleus?

3. How can we say that an atom is neutral when the outside parts of the atom are all negatively charged?

4. Since the outside parts of all atoms are negatively charged, how could one neutral atom be attracted to another neutral atom?

These questions can be assigned as written homework. Discussion in the next class period will be concentrated on the student responses to these questions.

NOTE: These questions will be quite difficult for most students to answer. If they are graded, the grading should be based on reasonable attempts instead of some strict criteria for the answers. The main purpose for assigning the questions is to encourage the students to spend some time giving thought to the questions prior to their discussion in class.

IV. THIRD DISCUSSION: APPLICATIONS OF CHARGE INTERACTIONS TO THE ATOMIC MODEL.

A. As stated at the end of the notes for the second discussion, this discussion should begin by having the students contribute their thoughts to the answering of the questions posed. The following items, numbered to correspond to the question numbers, give an idea of reasonable conclusions that can be drawn to answer the questions. It should be noted that the development of the ideas that answer these questions may require significant amounts of time and astute questioning on the part of the teacher. The questions should be dealt with in order and one at a time.

1. The forces holding the nucleus together must be of a different type than the observed forces acting between electrical charges. The presence of neutrons in the nucleus may play a role in holding the nucleus together. At any rate, these nuclear forces can be characterized as being:

   a. very short-range (nuclear fission can be used as an example of this);

   b. very strong (most nuclei are very difficult to break apart);

   c. independent of charge (the only charged particles in the nucleus are protons); and

   d. attractive (our model says that the nucleus <u>does</u> stay together).

2. Electrons must have enough total energy (potential and kinetic) to counteract the force of attraction so that they do not get pulled into the nucleus.

   a. To make this point, the teacher may wish to suspend pith balls on string from separate ring stands; induce opposite static charges on the balls; and then bring the ring stands close enough to each other to allow the balls to attract and touch. These balls can be called the "nucleus" and "electron".

   b. For simplicity, assume that all of the electron's energy is kinetic energy. Separate the two pith balls, transfer more charge to each one, and then start moving the "electron" back and forth near the "nucleus". If sufficient motion is given to the "electron", it will not be held in contact with the "nucleus" as it passes near.

3. The effect of a charge, as demonstrated with the electroscope, extends beyond the charge itself. The interaction of the positive and negative charge gives a neutral effect overall to the atom, <u>when</u> the atom contains equal numbers of protons and electrons.

4. Since the effect of a charge does extend beyond the charge itself, one could also hypothesize that atoms near each other could experience some attraction between the electrons of one atom and the protons of

a neighboring atom.

This discussion should be sufficient for completing the explanation of how the parts of an atom interact with each other. This explanation can be obtained using the model that the class has accepted (refer to the first discussion).

V. FOURTH DISCUSSION: ELECTRON ENERGY.

A. It has been established, in answering the questions from the second and third discussions, that electrons have enough energy to keep them from being drawn into the nucleus of the atom. The teacher will want to remain alert to the fact that electron energy is an <u>essential</u> part of our atomic model. This is true because of the explanations already allowed, but the importance of the concept becomes even more obvious as we lead up to the idea of electron energy levels.

B. A logical question to ask is, "What would happen if more energy were added to one of the electrons in a <u>neutral</u> atom?". (The electron would tend to move away from the atom because it would be able to overcome at least part of the attraction that holds it in the atom.)

1. Have the class review the comparative number of protons and electrons in a neutral atom (the numbers are equal).

2. Ask if the number of neutrons will affect the overall charge. Insist on supportive information. (No, because neutrons are not electrically charged.)

3. Return the discussion to the question about adding energy to an electron. Ask if enough energy could be added to remove an electron from an atom. (Yes, because there is a finite amount of positive charge attracting the electron. Given enough energy, therefore, the electron could be freed of the nuclear attraction.)

4. Invent <u>ion</u> --an atom that has an unequal number of positive and negative charges; that is, protons and electrons. A definition such as this should be written on the board. This definition can be easily expanded to cover groups of atoms with net charge, and this expansion will be made when the list of common ions is presented.

C. A series of questions follows from the previous reasoning and the invention of <u>ion</u>.

1. What is the electrical character of an atom that has had an electron removed from it? (It will have a charge of positive one.)

2. How does the energy of the removed electron compare to its energy when it was contained in the atom? (Its energy is greater than when it was contained.)

3. Could a neutral atom become a negatively charged ion? (yes, but only if the atom could pick up an electron. The rubber rod's negative charge can be used as evidence of such a possibility.)

4. How could the formation of a negative ion be explained in terms of our model of the atom? (The effect of a charge extends beyond the charge itself.)

D. During the discussion leading to the answers to Questions 1 and 2 above, two ideas should be stressed.

1. The nucleus of the atom is not affected in the formation of the ion. In fact, the charge on the nucleus (as indicated by the number of protons in the nucleus) is unique for a given element. Invent atomic number.

   a. Point out the place on the periodic table where the atomic number is listed. Have the students find the Z values on their own periodic tables, since different tables often list the numbers in different places.

   b. Briefly quiz the students about the atomic numbers of various elements to be sure that they know where to find the numbers.

2. Energy has to be added to an electron to remove it from an atom.

E. Answering Question 3 reinforces the idea that ion formation involves only a change in the number of electrons in an atom. Some students are tempted to say that a negative ion would be formed when an atom loses a proton. If the teacher questions these students about the implications of this reasoning, the students should be able to understand that any net charge of the atom will be due to changes in the number of electrons. For example:

1. When we discussed forming a positive ion, did we change the type of element that we were discussing? (No. The element remained the same. Only the net charge changed.)

2. Does our definition for an ion have anything to do with changing one element to another? (No.)

NOTE: Even if the students do not bring up the idea of changing the number of protons to form ions, the questioning suggested above should be initiated by the teacher.

F. Reading 6a: ATOMIC PUDDING, can be assigned at this time. This reading gives a brief overview of the historical development of atomic theory.

G. The effects of changing the numbers of protons and electrons should now be clear to the students. It is a good time now to ask how different numbers of neutrons in a nucleus would affect an atom. (The only effect would be that the atom would weigh more or less, depending on whether neutrons were "added or taken away".)

1. When students have discussed this question and have arrived at the idea above, invent <u>isotope</u> --atoms of the same element having different weights. Write the definition on the board.

2. Have the students explain their understandings of this definition. Stress ". . .of the same element. . ." and ask what will determine what the element is (<u>only</u> the number of protons).

3. Tell the students that the idea of isotopes can be used to explain the weight numbers on the periodic table; that is, the fact that most of them are not whole numbers. Most elements have two or more naturally occurring isotopes, and the <u>weighted average</u> of these can be assumed to be the weight listed on the table.

4. The idea of isotopes is expanded in Reading 6b: DOES THE AMOUNT OF MATTER MATTER?. After the students have had time to read this article, the teacher may wish to spend a class period in discussing some of the article's content and doing one or two <u>weighted average</u> calculations to show how table weights can be obtained. Suggested examples are boron with isotopes weighing 10 and 11, and

chlorine with isotopes weighing 35 and 37.

H. Question 4 from Section C of this discussion reinforces the idea that the effect of the charge on an object can be extended beyond that object. The teacher will have to allow time for the students to think about the work that has been done previously (that is, the field of force idea from the electroscope demonstration) before they can respond to this question. In fact, the teacher may have to provide clues to help the students relate previous work to the question. Pull out the electroscope again, if necessary.

I. It has now been established that ion formation, whether positive or negative, involves changes in the number of electrons associated with an atom. It is convenient, at this time, to return to the idea of adding energy to an electron to remove it from an atom.

1. Ask the class what has to happen for a neutral atom to become a positive ion. (Sufficient energy must be added to an electron to allow it to leave the attractive force field of the atom.)

2. Use their responses to invent <u>first ionization energy</u> -- the amount of energy required to remove one electron from a neutral atom. Write this on the board. Note: This definition does not specify that the atom is in the gaseous state to avoid any unnecessary confusion factor that may arise because of such a statement.

J. Point out to the class that the first ionization energies of elements have been determined experimentally, and present a listing of these first ionization energies for the first 18 elements on the periodic table. (This can be done by preparing an overlay of these values.)

NOTE: These data are taken from <u>Handbook of Chemistry and Physics,</u> Robert C. Weast, ed. The Chemical Rubber Co., 1967.

Atomic Number (Z)	Element Symbol	First Ionization Energy (Volts)
1	H	13.527
2	He	24.46
3	Li	5.6363
4	Be	9.28
5	B	8.257
6	C	11.217
7	N	14.48
8	O	13.55
9	F	17.34
10	Ne	21.47
11	Na	5.12
12	Mg	7.61
13	Al	5.96
14	Si	8.12
15	P	10.9
16	S	10.3
17	Cl	12.952
18	Ar	15.68

A brief description of the units employed is in order, and will probably be requested by the class. This can be done in terms of a general description of the procedure for determining ionization energies; that is, voltage is increased across a sample of substance until a current flows--electrons come off the atoms.

K. Given the first ionization energy data, we want to see if they can be of value to us in learning more about atoms. Ask the class how they might begin examining these data. They should have had experience with graphing data before, and this idea should come up.

1. Suggest that it may be easier to graph the data if the numbers were smaller.

2. The students have had valuable prior experience using ratios (comparisons) of numbers in laboratory activities. Apply the comparison ideas to the first ionization energy data by making ratios of the following ionization energies: hydrogen to hydrogen and helium to hydrogen. (Ratio values will be 1.00 and 1.81, respectively. These represent relative ionization energies.)

3. When these examples have been explained, present the class with a list of <u>relative first ionization energies.</u> This is a student handout sheet.

4. As a homework assignment, ask the students to graph the data, with atomic number (Z) on the x-axis and relative ionization energy (I) on the y-axis.

5. Tell them to connect each consecutive point with a straight line and then try to determine if patterns can be noted in these data. Have them write any patterns they see on the back side of their graph paper. Have them include any other observations about the graph that they can make.

6. In setting up the graph, the students should be advised to make the relative ionization energy axis at least <u>six</u> units high. (This arrangement will allow for the subsequent plotting of the second relative ionization energies, when that graphing is performed later.

VI. FIFTH DISCUSSION: PATTERNS IN FIRST IONIZATION ENERGIES

A. The first part of this discussion will be involved with recording the patterns perceived by the students after they have graphed the relative first ionization energy data.

B. List these patterns on the board. Encourage the students to include patterns that other students saw at the end of their own lists.

NOTE: You may find it valuable to obtain a list of these from the class "secretary".

C. Although the specific direction that the following discussion takes will depend on the list of patterns perceived, the following questions apply to ideas that should be developed in the discussion.

1. How do the patterns of the values relate to the arrangement of the elements on the periodic table? (The lower values for ionization energies correspond to elements on the left-hand side of the table, and the higher values correspond to elements on the right-hand side of the table. Values decrease as one moves down a column of elements.)

2. Which elements have consistently high values and which ones are consistently low? (I-A, II-A, III-A are consistently low and V-A through noble gases are consistently high.)

3. How can these values be interpreted in terms of how strongly an atom holds its electrons? (The higher the values, the more strongly the atom holds its electrons.)

4. Which elements would form positive ions the easiest? The hardest? (I-A elements and noble gases, respectively.)

5. How would you predict the first ionization energy of Cs to compare with that of Na? (This can be carried on to other families in the table. The first ionization energy of Cs should be lower.)

D. There are three important inventions that need to be made during this discussion. They are: <u>period of elements</u>, <u>family of elements</u>, and <u>periodic property</u>. Depending on the patterns noticed by the students and the discussion about those patterns, these inventions can be made to enter the discussion quite smoothly. It should be noted that ionzation energy is our only example, so far, of a periodic property.

E. At this point, the teacher should raise a question about how the amount of energy required to remove a second electron from an atom would compare to the amount of energy required to remove the first electron from a neutral atom. (The energy required for the second electron would be greater.)

   1. The students should be allowed to offer reasons why the second amount of energy is higher (for example, less repulsion among electrons giving the nucleus a greater attractive effect).

   2. Invent <u>second ionization energy,</u> and define it as the amount of energy required to remove one electron from a 1^+ ion.

F. Give the students copies of the second relative ionization energy data. The teacher will notice that these data are relative to the first ionization energy of hydrogen. This point should be made to the class. Obviously, there will be no data for hydrogen in this array. The teacher may want to ask the class why this is so. (Hydrogen has only one electron, when it is neutral.)

G. Instruct the students to make a graph from these data on the same paper they used for the first ionization energy graph. The graphing should be done in a manner similar to that of the previous graph; that is, consecutive points connected by straight lines.

H. This graphing should be done as a homework assignment. Ask the students to list patterns for this graph (and comparisons between it and the other graph) on the back side of the graph paper.

VII. SIXTH DISCUSSION:

PATTERNS IN SECOND IONIZATION ENERGIES

A. This discussion should begin by having the students volunteer the similarities and differences they observed between the graphs for first and second ionization energy. These could well be listed in columns on the board.

1. Of great importance in this analysis are the positions of the peaks and low points on the two graphs. The meanings of these positions will not be clear until later in the discussion, but noting their positions now is important.

B. Almost certainly some student will observe that the value for second ionization energy is always higher than that of first ionization energy for a given element.

1. Ask the class how much higher the second ionization energies are than the first ionization energies. Lead the discussion to the application of the ratio operation for making the comparisons.

2. Assign the computation of values for the ratio of second ionization energy to first ionization energy. Assign one calculation per student for helium through sodium and list the values obtained on the board. These values can be rounded off to the nearest whole number for simplicity of discussion.

3. Center the attention of the class on the comparison for helium. It may be helpful to use a diagram as an aid in the discussion. For example:

$$- 2+ -$$

a. Removal of the first electron (forming a 1+ ion) will take the amount of energy indicated on the first ionization energy graph. Ask the students to speculate about why the first ionization energy of helium is not quite twice the value for the first ionization energy of <u>hydrogen.</u> (Although the nuclear charge has doubled, the repulsion effect of the second electron must be such that the first ionization energy is lowered somewhat.)

b. It will be observed (from the value calculated for the second I.E. : first I.E. ratio) that roughly twice as much energy is required to remove the second electron from helium than was needed to remove the first electron. Ask the class to explain whether or not this is reasonable in terms of our model. (The class should conclude that it is reasonable.)

NOTE: The foregoing is based on an assumption that the two electrons in a neutral helium atom have equivalent energies. Since we have done nothing yet to try and specify electron energy values (either comparative or absolute), an assumption of equal energies for the two electrons is not unreasonable. We will return to this assumption a little later in the discussion.

c. Discuss in general the ratio values for the remaining elements (Z=3 through 11). The students will observe that all the values round to 2 or 3, except for lithium and sodium, whose values are <u>much</u> higher.

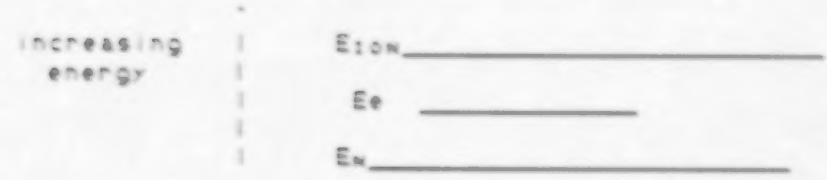
d. Ask if these values--excepting for lithium and sodium--can be explained in a way similar to that for lithium. (It would be reasonable to say that the explanations are similar.)

C. It should be clear that the ratio values obtained for lithium and sodium must be explained in terms other than an effective increase in attraction due to the loss of the first electron. By questioning the class, be sure that the students understand this idea.

1. Ask the class how the energies of the second electron removed from lithium and sodium compare with the energies of the second electron removed from the other eight elements under consideration. (These energies are much <u>lower</u>, as indicated by the much greater amount of energy that must be added to remove the second electron.)

a. It is important to review the essence of ionization energy again: that is, energy must be <u>added to</u> electrons to remove them from an atom.

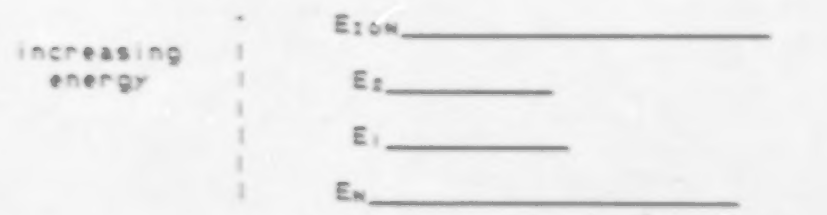
b. Review the explanation that the class has given for the majority of the ratios between second and first ionization energy. The following diagram should be helpful.



In this diagram, E_N is the energy an electron would have if it were next to the nucleus. Ee is the energy of two electrons in an atom--notice that this is consistent with the assumption of equal electron energies made earlier. E_{ion} is the amount of total energy needed for the electron to be free of the atom.

The explanation of this diagram can proceed in a way similar to the explanation for the first two ionizations of helium, and it will apply to the eight elements with small second to first ionization ratios. The teacher should be careful to insure that the students can explain why the second electron would require more energy, even though its initial energy is the same as that of the first electron removed.

D. Now the teacher can apply the same diagram to the explanation of the values for lithium and sodium. This, of course, will not work, and a diagram such as the following should result.



This shows the first electron coming off the atom from the higher energy E_2. The second electron to come off must have a lower initial energy, E_1.

E. Invent <u>electron energy level</u>. Note the interchangeable way that energy <u>level</u> and energy <u>amount</u> can be used. It is important that the students think of energy levels as amounts of electron energy.

1. Ask the students to refer to their graphs and ratio calculations to determine how many electrons can exist in the first (or lowest) energy level. (Work with the students to show that there can be a maximum of two electrons.)

NOTE: Some students have reasoned that the energies of these electrons would have to be higher than the energies of outer electrons because of the increase force of attraction on these inner electrons (the higher energy being required to keep these inner electrons from being pulled into the nucleus). To solve this problem, remind the students of our reference point for electron energy; that is, the amount of energy that must be added to remove it.

2. Ask the students to determine how many electrons could occupy the second energy level. (The students should be able to deduce that a maximum of eight electrons can be in the second energy level.)

   a. Ask if it is reasonable that there could be more electrons in the second than in the first energy level. (This is reasonable in terms of the three dimensionality of the model and in terms of the idea that the lower the energy level is, the closer it will be to the nucleus. Since electrons repel each other and effectively occupy a rather large space, the larger available space, at the second energy level distance, allows room for more electrons.)

3. The teacher can show the class a way to determine the <u>maximum</u> number of electrons that can occupy an energy level (that is, $2n^2$, where n is the number of the energy level). Work some examples.

NOTE: Some students may have noticed how the first two energy levels correspond to the first two periods on the periodic table. This fact may cause them some difficulty, when the computation above shows that 18 electrons can occupy the third energy level. The teacher should assure the class that this problem will be resolved a little later as we continue to study the data that we have.

F. Review the work done in this discussion by asking the class to interpret the meaning of the peaks on the two graphs of ionization energies. (If this question is restricted to the first two peaks on each of the graphs,

the peaks represent the energy required to remove an electron from a <u>full</u> first or second energy level.)

1. Ask the students to predict which elements would show peaks on graphs of third, fourth and fifth ionization energies. (The peaks would occur at beryllium and magnesium for third I.E.; boron and aluminum for fourth I.E.; and carbon and silicon for fifth I.E. Students are generally able to predict these readily.)

2. Have the students explain why their predictions would be valid.

B. READING 6c: HOW BIG IS AN ATOM? can serve well as a homework assignment at this time. Part of the assignment should be the graphing of the data and answering the questions at the end of the article. The ensuing class period should be involved with discussion of these graphs and the questions. Acceptable responses to these questions follow. (Note: The content of this article will be important in later discussions.)

NOTE: An overlay of the graph of these data would greatly facilitate the discussion of this graph.

1. Analysis of this graph will indicate that atomic radius is a periodic property for A-column elements. The list of trends and patterns will be similar in many ways to the lists made from the graphs of ionization energies. A noteworthy difference will be that the values generally <u>decrease</u> between peaks, instead of the general increase noted for ionization energies. The peaks on this graph will occur at the I-A elements.

2. This question will require considerable discussion. It is important to remind the students that these data refer to <u>atom sizes,</u> and they should not be confused with the energy data which were graphed earlier. There is an apparent contradiction in these data. Each successive element in a period has one more electron than the element preceding it, so it would seem (at first glance) that each successive element should have a larger radius because of greater electron repulsion. The opposite size trend is observed, however, with the exception of the noble gases.

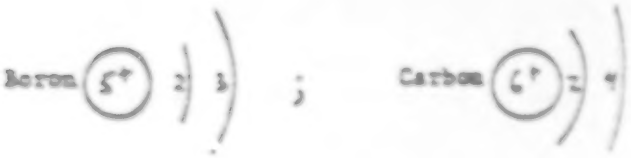
   a. Ask the students what factors are involved in determining the size of an atom. (The number of

protons and electrons and the distance between the nucleus and the highest energy electrons will determine atomic size.)

b. Ask what effect the number of protons, then the number of electrons (each taken separately) would have on atomic size. (The greater the number of protons, the greater the attraction for electrons. This should cause the size to decrease. The greater the number of electrons, the greater the repulsion. This repulsion should cause the size to increase.)

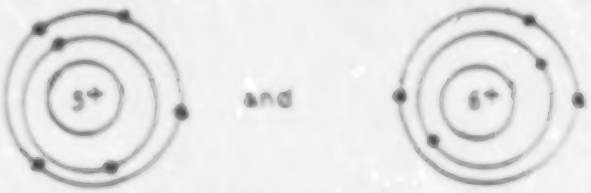
c. In light of the data, ask the class which factor is most effective when going from left to right in a period. (The nuclear charge; that is, the number of protons has the greater effect.)

d. Center the attention of the class on two successive elements (that is, boron and carbon). Ask the class to state how the electrons fit into energy levels for these two elements. As the students respond to this question, the teacher can supply a diagram, such as the following, to summarize the responses.



NOTE: This representation is used for organizing the electrons into energy levels and for providing a rough perspective for distance of energy levels from the nucleus. Be sure that the students understand the purpose and the limitations of this diagram (that is, we are not trying to specify exact distance of electrons from the nucleus, their positions with respect to each other in an energy level, or the way that the electrons move outside the nucleus.)

e. Point out to the class that the electrons that are farthest from the nucleus, for both of these atoms, are in the second energy level. Propose that we assume that the distance between a nucleus and the second energy level is the same for both of these atoms. With the aid of diagrams, such as



center the attention of the class on the attraction between the nucleus of boron and one of the electrons in the second energy level. Using our assumption of equal distance, ask the class how this attraction in boron would compare to the attraction between the carbon nucleus and one electron. (The attraction for the carbon electron should be greater. If this question poses a problem to the students, remind them of how the charge effect varied with distance and amount of charge during the electroscope demonstrations.)

f. Remind the students that, in reality, we cannot ignore the repulsion of other electrons (as we have in this argument), but the data support our explanation of decreasing atomic size.

g. Ask the class to offer an explanation for the great increase in size for the noble gas elements. (Probably the best response to this question is that with eight electrons in the highest energy level--or two, in the case of helium--the repulsive effect of the electrons takes precedence over the net attractive effect observed in the other elements.)

3. How would the atomic radii of lithium and beryllium compare to:

a. Each other? (Lithium is larger than beryllium.)

b. The rest of the second period radii? (The radii of lithium and beryllium atoms will be larger than the rest of the atoms of the second period.)

4. K, Ca, Ga, etc., should be larger than Na, Mg, Al, etc., because the trend from the second to the third period shows this pattern. For this question, the significant difference between periods two and three is that the second period has the second energy level as its highest energy level and the third period has the third energy level as its highest energy level. We could assume, therefore, that the highest energy level for K, Ca, Ga, etc., would be the fourth energy level.

NOTE: This idea will be expanded in the next discussion.

VIII. SEVENTH DISCUSSION:

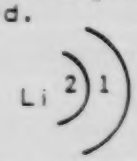
INTRODUCTION TO CHEMICAL NOMENCLATURE

A. Given a neutral atom, ask the class to state the energy level from which it would be easiest to remove an electron (the highest, or outermost, energy level).

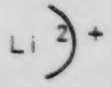
1. Have the class support their response in terms of the atomic model that has been developed.

2. Invent <u>valence level.</u> For purposes of definition, we will call the valence level the highest energy level that contains electrons, <u>when the atom is neutral.</u>

B. Ask the class to dictate the electron arrangement for a neutral lithium atom. When the students have found this atom and stated its electron arrangement, put the following diagram on the board.



1. Now ask the class to supply the electron arrangement for a lithium ion with a 1+ charge.



C. Repeat the questioning from Part B for sodium, and then for beryllium and boron--when all valence electrons have been removed from these three elements.

D. Ask the class what the electron arrangements for these ions resemble. (For all except the sodium ion, these ionic electron configurations resemble that of a neutral helium atom. The ionic electron configuration for sodium resembles a neutral neon atom.)

1. Continue this exercise by obtaining the ionic electron arrangements for a 2+ magnesium ion and a 3+ aluminum ion. (These ions will have electron configurations corresponding to neutral neon atoms.)

2. Ask the class if it would be reasonable to consider removing another electron from any of these ions. Require supporting arguments for their responses. (It would not be reasonable to remove another electron from these ions because of the comparatively great amount of energy that would be required to do so. It would be good to refer to the composite ionization energy overlay to reinforce this idea.)

E. Now that it has been established that these electron arrangements are stable, with reference to further ionization, invent noble gas electron configuration.

1. Because of the importance of this idea, it would be good to provide a written definition for the class, such as: A noble gas electron configuration is an electron arrangement that corresponds to that of a neutral noble gas atom.

F. Pose a question of the following type to the class: "Since a noble gas electron configuration seems to be a quite stable configuration for the electrons in an atom, what type of ion would you expect a chlorine atom to form?" (There are two responses that can be expected for this question; that is, the formation of a 7+ ion and/or a 1- ion. Both of these responses should be accepted, initially, and the second possibility should be obtained through questioning, if it is not proposed by the students. Draw diagrams of these ionic electron arrangements on the board.)

1. Ask the class to suggest whether one of these configurations is more reasonable than the other. (Using questioning, it should be established that the 1- ion would be more reasonable in terms of the amount of energy involved in forming the different ions. The idea of chlorine's relatively high attraction for electrons--as evidenced by its ionization energy--should be estabished as evidence

that it could also exert a relatively strong attraction on other electrons.)

2. Some students may suggest that the I-A, II-A and III-A elements could attain noble gas electron configurations by gaining electrons. This suggestion can be discussed in terms of the relatively low ionization energies (electron attractions) for these elements.

NOTE: This discussion of the I-A, II-A and III-A elements should be initiated by the teacher, if a student does not suggest it.

G. A good question now would be, "Where could a chlorine atom get the electron it needs to achieve the noble gas electron configuration?" (Students are usually eager to suggest that the needed electron could be obtained from another element, which leads to a question about which elements would be likely candidates to have electrons removed from them. These elements, of course, are ones with low ionization energies. Once this has been established, suggest sodium as an example element.)

1. Guide the students through a review of the idea that the relatively high ionization energy of chlorine (contrasted to the low ionization energy of sodium) means that the chlorine nucleus is exerting a substantial attractive force on its own valence electrons. This fact, combined with the responses they have given, implies that the nucleus of the chlorine atom could exert enough force on the valence electron of sodium to remove that electron from the metal and pull it into its own valence level.

2. Question the class about what electrical charge conditions exist after this electron transfer has occurred. The idea is that two ions of opposite charge now exist, both having noble gas electron configurations. The ions should then be strongly attracted to each other because of the force of attraction between opposite electrical charges.)

3. Ask the class if this process would be considered as a chemical change or a physical change (chemical).

H. Invent <u>ionic bonding.</u>

I. Using the ideas of noble gas electron configurations and ionic bonding, the teacher should lead the class in

establishing formulas for several other compounds each containing a positive and a negative ion from the A-columns (for example: aluminum chloride, sodium fluoride, potassium nitride, magnesium oxide, and other binary compounds of A-column elements the teacher may desire to use.)

NOTE: It will be necessary to rationalize the 2- and 3- charges on the oxide and nitride ions, respectively, but this should be a relatively simple matter in light of the previous discussion of ion formations. The teacher should point out that we have a model for the atom that allows us to explain the formulas for many compounds that we have used already.

J. Ask the class to dictate the electron arrangement for a neutral potassium atom. (The most common response to this question is that there are two in the first energy level, eight in the second, and nine in the third. This response is, of course, based on the fact that the third energy level can contain a maximum of eighteen electrons.)

1. Assuming the response above, ask if the ionization energy data support that arrangement.

   a. Have the students follow the basic trend in first ionization energy for the elements in the third period. Present this in such a manner that the students will supply the number of valence electrons in each successive element.

   b. At argon, there are eight valence electrons and a peak in ionization energy. In light of the trend in the period and the value for argon, ask the class if potassium's ionization energy should be high or low with the proposed electron arrangement. (It should be high--probably higher than that of argon.)

   c. Since the first ionization energy of potassium is very low, ask the class for an explanation of the data. (The most reasonable explanation is that there is at least one electron in the fourth energy level.)

NOTE: Some students will react to this idea because of the $2n^2$ rule, and the teacher will have to patiently point out that the data do not support the original arrangement. We are now attempting to provide an explanation that the data will support.

d. Ask how many electrons would be in the fourth energy level of potassium. (Only one. Allow the students to work on this before referring them to the second ionization energy graph, which shows that the second electron to be removed from potassium would have to come from a lower energy level.)

2. Ask the class to suggest the electron arrangement of a neutral calcium atom. (There are two possible answers to this question, as there were for potassium, but the students should now be more alert to the dictates of the data, as a result of the discussion of potassium. At any rate, the teacher should lead the class to the idea that calcium has two fourth-level valence electrons.)

K. Point out that the first row of B-column elements begins with element 21 and goes through element 30. From previous discussion it has been implied that the A-column family number corresponds to the number of valence electrons in an atom. This fact should be established clearly at this time by the teacher. If this rule applies to fourth period elements, then the electron added to each successive element in the B-columns cannot go in the valence level. Since there is room in the third energy level for these electrons, we can assume that they go into the third energy level.

L. It can now be shown that the period number corresponds to the number of the valence level for the elements in that period (that is, the valence level for the first period is the first energy level, and so forth).

NOTE: As a first approximation, we can assume that the B-column elements contain two valence electrons.

NOMENCLATURE EXERCISES

A. Hand out the blank ion worksheets and explain that we will list ions that occur commonly in chemical formulas.

NOTE: There are two main reasons for having the students fill out these lists of ions instead of handing them a prepared list: (1) The act of copying the ions

down will aid the students in learning them and (2) the groupings of ions by common properties and/or charge will make them take on added significance.

B. The teacher may wish to have the students supply charges for some metal ions or it may be deemed more desirable to ask what charges would be on ions of the I-A through III-A metals. Charges on monatomic ions should be related to the noble gas electron configurations repeatedly during this presentation. This idea also applies to the presentation of monatomic anions. The teacher may then expose the ions of I-A, II-A and III-A metals so that the students may write them on their sheets. (It may be convenient for the teacher to fill out the entire table and make a transparency of it. Different sections of the transparency can be uncovered, as they are applicable to the discussion.)

C. Students will probably want to know why hydrogen was not included with the I-A metals, since many periodic tables list it with the I-A elements. It would be good to discuss this question.

1. Hydrogen most commonly forms a (1+) ion, but it can, and does, form a (1-) ion.

2. The (1-) ion has the noble gas configuration of helium, which is reasonable with respect to other ions.

3. The (1+) ion has no electrons, so its configuration differs, in that regard, from the I-A elements.

4. It has been stated that hydrogen behaves as a family of its own, having ion characteristics like both the I-A and VII-A elements, but with an intermediate value for ionization energy. It is a gas at room temperature.

D. After the aluminum ion has been written down, the B-column ions may be introduced. The point about the multiple possibilities for the charges of B-column ions should be made.

1. When the names of these ions are written, the appropriate Roman numeral should be included in parentheses after the name. <u>Stress the fact that these Roman numerals appear only in the written name, not in the formula.</u>

2. A few of the B-column ions listed will not have Roman numerals after them. The reason for this is that the other charges are so uncommon that their listing is unwarranted in a listing of common ions or that they have only one ionic charge (for example, Ag^+).

3. Two of the ions listed with the B-column ions (Pb and Sn) are actually in the IV-A column. For now these can be treated as exceptions to the general rules for A-column ion charges.

4. Special note should be made of the mercury (I) ion formula and its use in compound formulas.

5. If the teacher so desires, a brief summary of the older, less-used names for the B-column ions may prove interesting to most students.

6. The uniqueness of the ammonium ion should be noted--not only is it polyatomic, but it is made up of two non-metals. The combination, however, has a positive charge.

E. The negative ions (anions) can be listed in several ways that would aid in learning them. The method of listing them given here provides for the utilization of element position on the periodic table for the monatomic anions. The polyatomic ions (as much as is possible) will be grouped in such a way that common characteristics can be used to help the students learn the ions.

1. The VII-A anions, through iodide, can be listed first. Spellings should be emphasized. The teacher should present the generalization concerning the naming of binary (two element) compounds (that is, the first element is named and then the ending of the name of the second element is dropped and replaced by "-ide").

2. The doubly-negative oxide and sulfide ions may be exposed next, as the common ions representing the VI-A column.

3. One ion (the nitride ion) is listed from the V-A column. The reason for presenting this ion is not that it is so common, but its listing allows for extension of the concept of ion formation in terms of noble gas electron configuration attainment.

NOTE: Someone may raise a question about the charge that would be formed by a IV-A element. It should be sufficient, for the time being, to state that there are no common ions of IV-A elements. In terms of the present discussion, it would seem equally logical for a IV-A element to form a (4+) or a (4-) ion. The ideas involved with covalent bonding will be discussed at a later time (see Investigation 17), and it would be in the best interest of most students' understanding to avoid confusing this presentation by adding the concept of covalent bonding.

F. Polyatomic anions may be presented next. Some of these ions will have to be memorized. Others can be learned by learning relationships to other ions with similar characteristics. An example presentation method for one group of ions with similar characteristics follows (for polyatomic ions containing chlorine and oxygen).

1. The reference ion is the one that occurs most commonly--chlorate. It should be memorized.

2. All the other ions in the group differ only in the number of oxygen atoms contained in them. The charges remain the same.

3. Prefixes and suffixes to the root name (-chlor-) designate the number of oxygen atoms contained.

   a. ChlorITE contains one less oxygen than chlorATE.

   b. HYPOchlorITE contains one less oxygen than chlorITE.

   c. PERchlorATE contains one more oxygen than chlorATE.

NOTE: The teacher might point out again that the binary ion for this group has the same charge, but ends in "ide".

G. Listing of sulfate and sulfite, nitrate and nitrite should follow. In relation to these, sulfide has a 2- charge, but nitride has a 3- charge.

H. Phosphate is listed next followed by the monohydrogen and dihydrogen phosphate ions. The charge effect from adding hydrogen to these ions should be noted and an explanation for it should be solicited from the class. (The positive hydrogen cancels a negative charge.)

I. The remaining hydrogen-containing anions can then be revealed.

J. Ions with the "thio-" prefix are presented next. The only two listed are thiosulfate and thiocyanate. It can be noted that the "thio-" prefix indicates that a sulfur atom has replaced an oxygen atom in an ion.

K. Two ions, which contain more than one element but end in "-ide" are the cyanide ion and the hydroxide ion. These exceptions to previously established general rules should be emphasized.

L. The remaining anions seem to lack any clear means of relationship to other ions listed.

M. ASSIGNMENT: Ten compound sheets are available to allow students to practice writing compounds from the list of common ions. It is advisable to help the students get started with the first sheet. In this way, the teacher can be sure that the students understand the proper placement of parentheses and subscripts in the formulas for various compounds.

COMMON POSITIVE IONS (CATIONS)

NAME	SYMBOL (with charge)	NAME	SYMBOL (with charge)
I-A		**B-COLUMN CATIONS (cont'd)**	
lithium	Li^{+}	iron (III)	Fe^{3+}
sodium	Na^{+}	cobalt (II)	Co^{2+}
potassium	K^{+}	cobalt (III)	Co^{3+}
cesium	Cs^{+}	nickel (II)	Ni^{2+}
hydrogen	H^{+}	nickel (III)	Ni^{3+}
		copper (I)	Cu^{+}
II-A		copper (II)	Cu^{2+}
magnesium	Mg^{2+}	zinc	Zn^{2+}
calcium	Ca^{2+}	silver	Ag^{+}
barium	Ba^{2+}	tin (II)	Sn^{2+}
III-A		tin (IV)	Sn^{4+}
aluminum	Al^{3+}	mercury (I)	Hg_2^{2+}
		mercury (II)	Hg^{2+}
B-COLUMN CATIONS		lead	Pb^{2+}
chromium (II)	Cr^{2+}	*ammonium	NH_4^{+}
chromium (III)	Cr^{3+}		
iron (II)	Fe^{2+}		

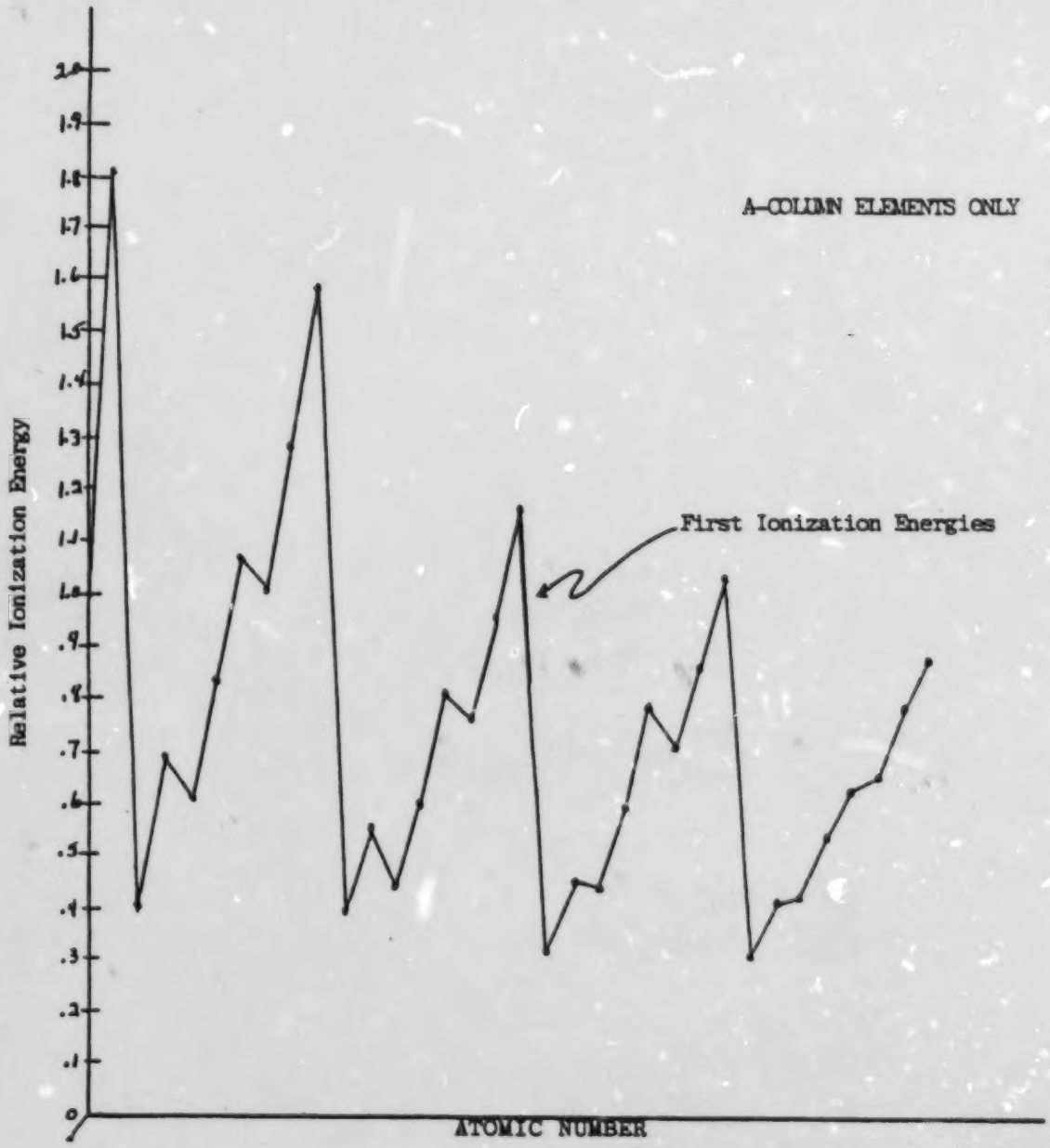
*polyatomic cations

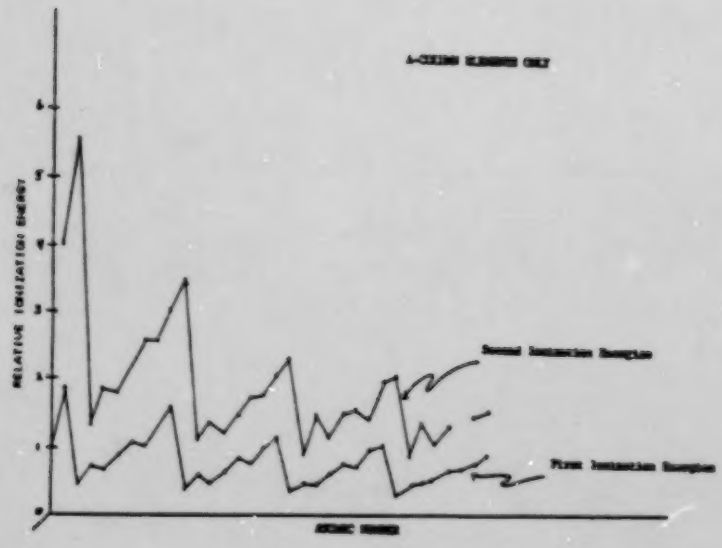
COMMON NEGATIVE IONS (ANIONS)

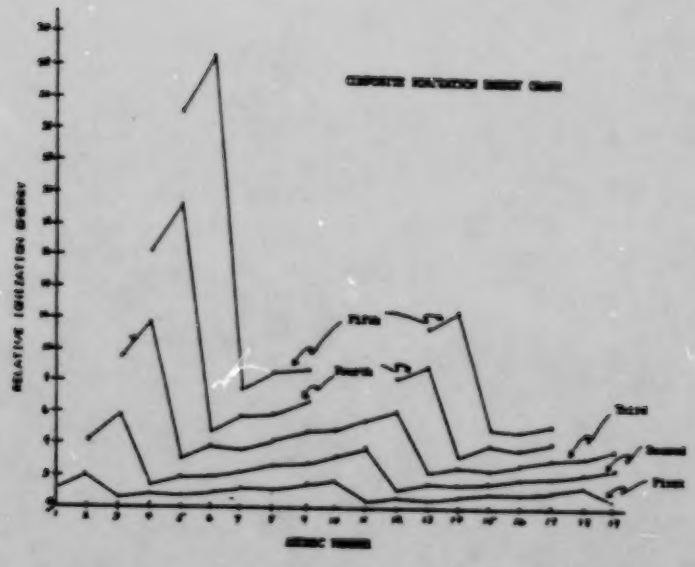
NAME	SYMBOL (with charge)	NAME	SYMBOL (with charge)
VII-A		**POLYATOMIC ANIONS (cont'd)**	
fluoride	F^-	nitrite	NO_2^-
chloride	Cl^-	phosphate	PO_4^{3-}
bromide	Br^-	monohydrogen phosphate	HPO_4^{2-}
iodide	I^-	dihydrogen phosphate	$H_2PO_4^-$
VI-A		carbonate	CO_3^{2-}
oxide	O^{2-}	bicarbonate	HCO_3^-
sulfide	S^{2-}	bisulfate	HSO_4^-
V-A		bisulfite	HSO_3^-
nitride	N^{3-}	bisulfide	HS^-
POLYATOMIC ANIONS		thiosulfate	$S_2O_3^{2-}$
chlorate	ClO_3^-	thiocyanate	SCN^-
chlorite	ClO_2^-	cyanide	CN^-
hypochlorite	ClO^-	hydroxide	OH^-
perchlorate	ClO_4^-	acetate	$C_2H_3O_2^-$
sulfate	SO_4^{2-}	chromate	CrO_4^{2-}
sulfite	SO_3^{2-}	dichromate	$Cr_2O_7^{2-}$
nitrate	NO_3^-	oxalate	$C_2O_4^{2-}$
		permanganate	MnO_4^-

INVESTIGATION 6: ATOMIC THEORY AND MATTER

RELATIVE IONIZATION ENERGY



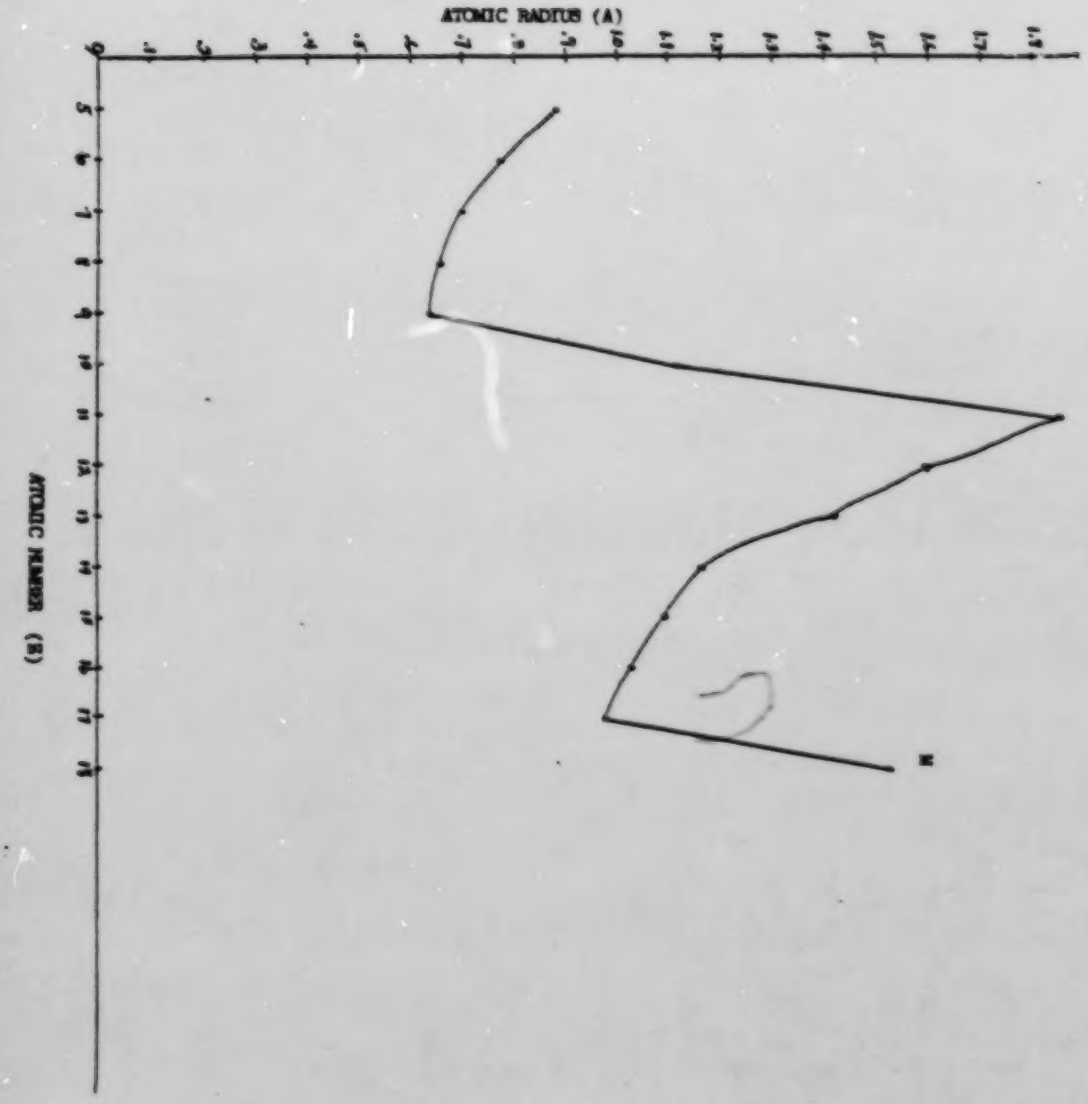




BEST COPY AVAILABLE

INVESTIGATION 6: ATOMIC THEORY AND MATTER

ATOMIC RADIUS VS. ATOMIC NUMBER



BEST COPY AVAILABLE

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 1

Binary Compounds with A-Column Cations (Part 1)

A. Hydrogen

1. hydrogen fluoride (HF)
2. hydrogen chloride (HCl)
3. hydrogen bromide (HBr)
4. hydrogen iodide (HI)
5. hydrogen oxide (H_2O)
6. hydrogen sulfide (H_2S)
7. hydrogen nitride (H_3N)

B. Lithium

1. lithium fluoride (LiF)
2. lithium chloride (LiCl)
3. lithium bromide (LiBr)
4. lithium iodide (LiI)
5. lithium oxide (Li_2O)
6. lithium sulfide (Li_2S)
7. lithium nitride (Li_3N)

C. Sodium

1. sodium fluoride (NaF)
2. sodium chloride (NaCl)
3. sodium bromide (NaBr)
4. sodium iodide (NaI)
5. sodium oxide (Na_2O)
6. sodium sulfide (Na_2S)
7. sodium nitride (Na_3N)

D. Potassium

1. potassium fluoride (KF)
2. potassium chloride (KCl)
3. potassium bromide (KBr)
4. potassium iodide (KI)
5. potassium oxide (K_2O)
6. potassium sulfide (K_2S)
7. potassium nitride (K_3N)

C. Cesium

1. cesium fluoride (CsF)
2. cesium chloride (CsCl)
3. cesium bromide (CsBr)
4. cesium iodide (CsI)
5. cesium oxide (Cs_2O)
6. cesium sulfide (Cs_2S)
7. cesium nitride (Cs_3N)

F. Aluminum

1. aluminum fluoride (AlF_3)
2. aluminum chloride ($AlCl_3$)
3. aluminum bromide ($AlBr_3$)
4. aluminum iodide (AlI_3)
5. aluminum oxide (Al_2O_3)
6. aluminum sulfide (Al_2S_3)
7. aluminum nitride (AlN)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 2

Binary Compounds With A-Column Cations (Part 2)

A. Magnesium

1. magnesium fluoride (MgF_2)
2. magnesium chloride ($MgCl_2$)
3. magnesium bromide ($MgBr_2$)
4. magnesium iodide (MgI_2)
5. magnesium oxide (MgO)
6. magnesium sulfide (MgS)
7. magnesium nitride (Mg_3N_2)

B. Calcium

1. calcium fluoride (CaF_2)
2. calcium chloride ($CaCl_2$)
3. calcium bromide ($CaBr_2$)
4. calcium iodide (CaI_2)
5. calcium oxide (CaO)
6. calcium sulfide (CaS)
7. calcium nitride (Ca_3N_2)

C. Barium

1. barium fluoride (BaF_2)
2. barium chloride ($BaCl_2$)
3. barium bromide ($BaBr_2$)
4. barium iodide (BaI_2)
5. barium oxide (BaO)
6. barium sulfide (BaS)
7. barium nitride (Ba_3N_2)

D. Test Your Understanding

1. boron fluoride (BF_3)
2. sodium phosphide (Na_3P)
3. calcium arsenide (Ca_3As_2)
4. hydrogen telluride (H_2Te)
5. rubidium selenide (Rb_2Se)
6. beryllium bromide ($BeBr_2$)
7. strontium oxide (SrO)
8. gallium chloride ($GaCl_3$)
9. magnesium phosphide (Mg_3P_2)
10. potassium selenide (K_2Se)

E. Write the Names

1. H_2S (hydrogen sulfide)
2. LiBr (lithium bromide)
3. NaF (sodium fluoride)
4. KI (potassium iodide)
5. CsCl (cesium chloride)
6. Al_2S_3 (aluminum sulfide)
7. Mg_3N_2 (magnesium nitride)
8. MgO (magnesium oxide)
9. $CaCl_2$ (calcium chloride)
10. BaS (barium sulfide)

COMPOUND SHEET 3

Binary Compounds With B-Column Cations
(elements with only one common charge)

A. Zinc

1. zinc fluoride (ZnF_2)
2. zinc chloride ($ZnCl_2$)
3. zinc bromide ($ZnBr_2$)
4. zinc iodide (ZnI_2)
5. zinc oxide (ZnO)
6. zinc sulfide (ZnS)
7. zinc nitride (Zn_3N_2)

B. Silver

1. silver fluoride (AgF)
2. silver chloride (AgCl)
3. silver bromide (AgBr)
4. silver iodide (AgI)
5. silver oxide (Ag_2O)
6. silver sulfide (Ag_2S)
7. silver nitride (Ag_3N)

C. Lead

1. lead fluoride (PbF_2)
2. lead chloride ($PbCl_2$)
3. lead bromide ($PbBr_2$)
4. lead iodide (PbI_2)
5. lead oxide (PbO)
6. lead sulfide (PbS)
7. lead nitride (Pb_3N_2)

D. Test Your Understanding

1. zinc phosphide (Zn_3P_2)
2. zinc telluride (ZnTe)
3. silver arsenide (Ag_3As)
4. silver selenide (Ag_2Se)
5. lead arsenide (Pb_3As_2)
6. lead telluride (PbTe)

E. Write the Names

1. ZnO (zinc oxide)
2. ZnS (zinc sulfide)
3. AgCl (silver chloride)
4. Ag_2O (silver oxide)
5. Pb_3N_2 (lead nitride)
6. PbI_2 (lead iodide)
7. AgI (silver iodide)
8. Ag_3N (silver nitride)
9. $ZnCl_2$ (zinc chloride)
10. PbO (lead oxide)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 4

Binary Compounds With B-Column Cations
(elements with more than one
common charge: Part 1)

A. Chromium (II)

1. chromium (II) fluoride (CrF_2)
2. chromium (II) chloride ($CrCl_2$)
3. chromium (II) oxide (CrO)
4. chromium (II) sulfide (CrS)
5. chromium (II) nitride (Cr_3N_2)

B. Chromium (III)

1. chromium (III) bromide ($CrBr_3$)
2. chromium (III) iodide (CrI_3)
3. chromium (III) oxide (Cr_2O_3)
4. chromium (III) sulfide (Cr_2S_3)
5. chromium (III) nitride (CrN)

C. Iron (II)

1. iron (II) fluoride (FeF_2)
2. iron (II) chloride ($FeCl_2$)
3. iron (II) oxide (FeO)
4. iron (II) sulfide (FeS)
5. iron (II) nitride (Fe_3N_2)

D. Iron (III)

1. iron (III) bromide ($FeBr_3$)
2. iron (III) iodide (FeI_3)
3. iron (III) oxide (Fe_2O_3)
4. iron (III) sulfide (Fe_2S_3)
5. iron (III) nitride (FeN)

E. Cobalt (II)

1. cobalt (II) fluoride (CoF_2)
2. cobalt (II) chloride ($CoCl_2$)
3. cobalt (II) oxide (CoO)
4. cobalt (II) sulfide (CoS)
5. cobalt (II) nitride (Co_3N_2)

F. Cobalt (III)

1. cobalt (III) bromide ($CoBr_3$)
2. cobalt (III) iodide (CoI_3)
3. cobalt (III) oxide (Co_2O_3)
4. cobalt (III) sulfide (Co_2S_3)
5. cobalt (III) nitride (CoN)

G. Nickel (II)

1. nickel (II) fluoride (NiF_2)
2. nickel (II) chloride ($NiCl_2$)
3. nickel (II) oxide (NiO)
4. nickel (II) sulfide (NiS)
5. nickel (II) nitride (Ni_3N_2)

H. Nickel (III)

1. nickel (III) bromide ($NiBr_3$)
2. nickel (III) iodide (NiI_3)
3. nickel (III) oxide (Ni_2O_3)
4. nickel (III) sulfide (Ni_2S_3)
5. nickel (III) nitride (NiN)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 5

Binary Compounds with B-Column Cations
(elements with more than one
common charge: Part 2)

A. Copper (I)

1. copper (I) fluoride (CuF)
2. copper (I) chloride (CuCl)
3. copper (I) oxide (Cu_2O)
4. copper (I) sulfide (Cu_2S)
5. copper (I) nitride (Cu_3N)

B. Copper (II)

1. copper (II) bromide ($CuBr_2$)
2. copper (II) iodide (CuI_2)
3. copper (II) oxide (CuO)
4. copper (II) sulfide (CuS)
5. copper (II) nitride (Cu_3N_2)

C. Tin (II)

1. tin (II) fluoride (SnF_2)
2. tin (II) chloride ($SnCl_2$)
3. tin (II) oxide (SnO)
4. tin (II) sulfide (SnS)
5. tin (II) nitride (Sn_3N2)

D. Tin (IV)

1. tin (IV) bromide ($SnBr_4$)
2. tin (IV) iodide (SnI_4)
3. tin (IV) oxide (SnO_2)
4. tin (IV) sulfide (SnS_2)
5. tin (IV) nitride (Sn_3N_4)

E. Mercury (I)

1. mercury (I) fluoride (Hg_2F_2)
2. mercury (I) chloride (Hg_2Cl_2)
3. mercury (I) oxide (Hg_2O)
4. mercury (I) sulfide (Hg_2S)

F. Mercury (II)

1. mercury (II) bromide ($HgBr_2$)
2. mercury (II) iodide (HgI_2)
3. mercury (II) oxide (HgO)
4. mercury (II) sulfide (HgS)
5. mercury (II) nitride (Hg_3N_2)

6. Write the Names

1. CrO (chromium II oxide)
2. Cr_2O_3 (chromium III oxide)
3. $CoCl_2$ (cobalt II chloride)
4. $CuCl_2$ (copper II chloride)
5. Cu_2S (copper I sulfide)
6. SnF_2 (tin II fluoride)
7. SnO_2 (tin IV oxide)
8. FeO (iron II oxide)
9. Fe_2S_3 (iron III sulfide)
10. $NiCl_2$ (nickel II chloride)
11. Hg_2Cl_2 (mercury I chloride)
12. $HgCl_2$ (mercury II chloride)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 6
Review of Binary Compounds

A. Write the Formulas

1. hydrogen chloride (HCl)
2. hydrogen telluride (H_2Te)
3. lithium oxide (Li_2O)
4. sodium bromide ($NaBr$)
5. potassium chloride (KCl)
6. cesium iodide (CsI)
7. aluminum oxide (Al_2O_3)
8. magnesium nitride (Mg_3N_2)
9. calcium oxide (CaO)
10. barium chloride ($BaCl_2$)
11. zinc sulfide (ZnS)
12. silver sulfide (Ag_2S)
13. lead iodide (PbI_2)
14. lead oxide (PbO)
15. chromium (III) oxide (Cr_2O_3)
16. iron (II) sulfide (FeS)
17. iron (III) oxide (Fe_2O_3)
18. cobalt (III) nitride (CoN)
19. nickel (II) sulfide (NiS)
20. nickel (III) bromide ($NiBr_3$)
21. mercury (II) sulfide (HgS)
22. mercury (I) iodide (Hg_2I_2)
23. copper (I) chloride ($CuCl$)
24. tin (IV) chloride ($SnCl_4$)
25. copper (II) nitride (Cu_3N_2)

B. Write the Names

1. Hg_2Cl_2 (mercury I chloride)
2. CaS (calcium sulfide)
3. $LiBr$ (lithium bromide)
4. ZnO (zinc oxide)
5. Na_2O (sodium oxide)
6. BaI_2 (barium iodide)
7. NiF_2 (nickel II fluoride)
8. SnO_2 (tin IV oxide)
9. $CoCl_2$ (cobalt II chloride)
10. AlN (aluminum nitride)
11. Fe_2S_3 (iron III sulfide)
12. Na_2S (sodium sulide)
13. Na_3N (sodium nitride)
14. AgI (silver iodide)
15. PbS (lead sulfide)
16. HF (hydrogen fluoride)
17. CsF (cesium fluoride)
18. SnF_2 (tin II fluoride)
19. HgI_2 (mercury II iodide)
20. CuO (copper II oxide)
21. H_2S (hydrogen sulfide)
22. MgO (magnesium oxide)
23. Al_2S_3 (aluminum sulfide)
24. CrF_2 (chromium II fluoride)
25. Cs_2O (cesium oxide)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 7

Compounds Containing Anions in the Chlorate, Sulfate, and Nitrate Series

A. Chlorate

1. potassium chlorate ($KClO_3$)
2. magnesium chlorate [$Mg(ClO_3)_2$]
3. nickel (II) chlorate [$Ni(ClO_3)_2$]
4. ammonium chlorate (NH_4ClO_3)
5. aluminum chlorate [$Al(ClO_3)_3$]

B. Chlorite

1. lithium chlorite ($LiClO_2$)
2. chromium (II) chlorite [$Cr(ClO_2)_2$]
3. cobalt (III) chlorite [$Co(ClO_2)_3$]
4. mercury (II) chlorite [$Hg(ClO_2)_2$]
5. zinc chlorite [$Zn(ClO_2)_2$]

C. Hypochlorite

1. sodium hypochlorite ($NaClO$)
2. barium hypochlorite [$Ba(ClO)_2$]
3. silver hypochlorite ($AgClO$)
4. copper (I) hypochlorite ($CuClO$)
5. cesium hypochlorite ($CsClO$)

D. Perchlorate

1. calcium perchlorate [$Ca(ClO_4)_2$]
2. nickel (III) perchlorate [$Ni(ClO_4)_3$]
3. lithium perchlorate ($LiClO_4$)
4. tin (IV) perchlorate [$Sn(ClO_4)_4$]

E. Sulfate

1. sodium sulfate (Na_2SO_4)
2. barium sulfate ($BaSO_4$)
3. hydrogen sulfate (H_2SO_4)
4. iron (III) sulfate [$Fe_2(SO_4)_3$]
5. ammonium sulfate [$(NH_4)_2SO_4$]

F. Sulfite

1. silver sulfite (Ag_2SO_3)
2. magnesium sulfite ($MgSO_3$)
3. tin (II) sulfite ($SnSO_3$)
4. aluminum sulfite [$Al_2(SO_3)_3$]
5. iron (II) sulfite ($FeSO_3$)

G. Nitrate

1. sodium nitrate ($NaNO_3$)
2. silver nitrate ($AgNO_3$)
3. copper (II) nitrate [$Cu(NO_3)_2$]
4. chromium (III) nitrate [$Cr(NO_3)_3$]
5. lead nitrate [$Pb(NO_3)_2$]

H. Nitrite

1. potassium nitrite (KNO_2)
2. aluminum nitrite [$Al(NO_2)_3$]
3. calcium nitrite [$Ca(NO_2)_2$]
4. cobalt (II) nitrite [$Co(NO_2)_2$]

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 8

Compounds with Polyatomic Anions that Contain Hydrogen

A. Monohydrogen Phosphate

1.	lithium monohydrogen phosphate	Li_2HPO_4
2.	calcium monohydrogen phosphate	$CaHPO_4$
3.	nickel (III) monohydrogen phosphate	$Ni_2(HPO_4)_3$
4.	mercury (II) monohydrogen phosphate	$HgHPO_4$
5.	copper (I) monohydrogen phosphate	Cu_2HPO_4

B. Dihydrogen Phosphate

1.	iron (III) dihydrogen phosphate	$Fe(H_2PO_4)_3$
2.	potassium dihydrogen phosphate	KH_2PO_4
3.	lead dihydrogen phosphate	$Pb(H_2PO_4)_2$
4.	magnesium dihydrogen phosphate	$Mg(H_2PO_4)_2$
5.	lithium dihydrogen phosphate	LiH_2PO_4

C. Bicarbonate

1.	sodium bicarbonate	$NaHCO_3$
2.	copper (II) bicarbonate	$Cu(HCO_3)_2$
3.	silver bicarbonate	$AgHCO_3$
4.	nickel (II) bicarbonate	$Ni(HCO_3)_2$
5.	calcium bicarbonate	$Ca(HCO_3)_2$

D. Bisulfate

1.	chromium (II) bisulfate	$Cr(HSO_4)_2$
2.	aluminum bisulfate	$Al(HSO_4)_3$
3.	zinc sulfate	$Zn(HSO_4)_2$
4.	mercury (I) bisulfate	$Hg_2(HSO_4)_2$
5.	potassium bisulfate	$KHSO_4$

E. Bisulfite

1.	nickel (II) bisulfite	$Ni(HSO_3)_2$
2.	potassium bisulfite	$KHSO_3$
3.	chromium (III) bisulfite	$Cr(HSO_3)_3$
4.	copper (I) bisulfite	$CuHSO_3$
5.	barium bisulfite	$Ba(HSO_3)_2$

F. Bisulfide

1.	zinc bisulfide	$Zn(HS)_2$
2.	cesium bisulfide	$CsHS$
3.	iron (III) bisulfide	$Fe(HS)_3$
4.	cobalt (II) bisulfide	$Co(HS)_2$
5.	silver bisulfide	$AgHS$

G. Hydroxide

1.	barium hydroxide	$Ba(OH)_2$
2.	sodium hydroxide	$NaOH$
3.	tin (IV) hydroxide	$Sn(OH)_4$
4.	chromium (III) hydroxide	$Cr(OH)_3$
5.	calcium hydroxide	$Ca(OH)_2$

H. Acetate

1.	copper (II) acetate	$Cu(C_2H_3O_2)_2$
2.	hydrogen acetate	$HC_2H_3O_2$
3.	ammonium acetate	$NH_4C_2H_3O_2$
4.	chromium (III) acetate	$Cr(C_2H_3O_2)_3$
5.	silver acetate	$AgC_2H_3O_2$

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 9

Compounds Containing Phosphate, Carbonate, Thiosulfate, Thiocyanate, Cyanide, Chromate, Dichromate, Oxalate, and Permanganate

A. Phosphate

1.	potassium phosphate	K_3PO_4
2.	magnesium phosphate	$Mg_3(PO_4)_2$
3.	aluminum phosphate	$AlPO_4$
4.	sodium phosphate	Na_3PO_4

B. Carbonate

1.	sodium carbonate	Na_2CO_3
2.	calcium carbonate	$CaCO_3$
3.	copper (II) carbonate	$CuCO_3$
4.	chromium (III) carbonate	$Cr_2(CO_3)_3$

C. Thiosulfate

1.	lithium thiosulfate	$Li_2S_2O_3$
2.	silver thiosulfate	$Ag_2S_2O_3$
3.	zinc thiosulfate	ZnS_2O_3
4.	cobalt (III) thiosulfate	$Co_2(S_2O_3)_3$

D. Thiocyanate

1.	mercury (II) thiocyanate	$Hg(SCN)_2$
2.	copper (I) thiocyanate	$CuSCN$
3.	tin (IV) thiocyanate	$Sn(SCN)_4$
4.	potassium thiocyanate	$KSCN$

E. Cyanide

1.	silver cyanide	$AgCN$
2.	aluminum cyanide	$Al(CN)_3$
3.	barium cyanide	$Ba(CN)_2$
4.	copper (I) cyanide	$CuCN$

F. Chromate

1.	zinc chromate	$ZnCrO_4$
2.	iron (III) chromate	$Fe_2(CrO_4)_3$
3.	sodium chromate	Na_2CrO_4
4.	lead chromate	$PbCrO_4$

G. Dichromate

1.	magnesium dichromate	$MgCr_2O_7$
2.	potassium dichromate	$K_2Cr_2O_7$
3.	cobalt (III) dichromate	$Co_2(Cr_2O_7)_3$
4.	tin (II) dichromate	$SnCr_2O_7$

H. Oxalate

1.	copper (I) oxalate	$Cu_2C_2O_4$
2.	hydrogen oxalate	$H_2C_2O_4$
3.	barium oxalate	BaC_2O_4
4.	nickel (III) oxalate	$Ni_2(C_2O_4)_3$

I. Permanganate

1.	potassium permanganate	$KMnO_4$
2.	iron (II) permanganate	$Fe(MnO_4)_2$
3.	chromium (III) permanganate	$Cr(MnO_4)_3$
4.	copper (I) permanganate	$CuMnO_4$

E. Cyanide

1.	silver cyanide	$AgCN$
2.	aluminum cyanide	$Al(CN)_3$
3.	barium cyanide	$Ba(CN)_2$
4.	copper (I) cyanide	$CuCN$

INVESTIGATION 6: ATOMIC THEORY AND MATTER

COMPOUND SHEET 10

Review of Compounds Containing Polyatomic Anions

A. Write the Names

1.	$KClO_3$	potassium chlorate
2.	$LiClO_2$	lithium chlorite
3.	$NaClO$	sodium hypochlorite
4.	$Ca(ClO_4)_2$	calcium perchlorate
5.	$BaSO_4$	barium sulfate
6.	$Al_2(SO_3)_3$	aluminum sulfite
7.	$Pb(NO_3)_2$	lead nitrate
8.	$Ca(NO_2)_2$	calcium nitrate
9.	$HgHPO_4$	mercury II monohydrogen phosphate
10.	KH_2PO_4	potassium dihydrogen phosphate
11.	$NaHCO_3$	sodium bicarbonate
12.	$Zn(HSO_4)_2$	zinc bisulfate
13.	$Cr(HSO_3)_3$	chromium III bisulfite
14.	$CsHS$	cesium bisulfide
15.	$Ca(OH)_2$	calcium hydroxide
16.	$NH_4C_2H_3O_2$	ammonium acetate
17.	$Mg_3(PO_4)_2$	magnesium phosphate
18.	$CaCO_3$	calcium carbonate
19.	$Ag_2S_2O_3$	silver thiosulfate
20.	$KSCN$	potassium thiocyanate
21.	$Ba(CN)_2$	barium cyanide
22.	$ZnCrO_4$	zinc chromate
23.	$SnCr_2O_7$	tin II dichromate
24.	$Cu_2C_2O_4$	copper I oxalate
25.	$KMnO_4$	potassium permanganate

COMPOUND SHEET 10 (Continued)

B. Write the Formulas

1.	chromium (III) permanganate	$Cr(MnO_4)_3$
2.	aluminum cyanide	$Al(CN)_3$
3.	sodium phosphate	Na_3PO_4
4.	copper (II) carbonate	$CuCO_3$
5.	potassium dichromate	$K_2Cr_2O_7$
6.	barium oxalate	BaC_2O_4
7.	copper (I) thiocyanate	$CuSCN$
8.	lead chromate	$PbCrO_4$
9.	zinc thiosulfate	ZnS_2O_3
10.	ammonium chlorate	NH_4ClO_3
11.	silver sulfite	Ag_2SO_3
12.	cesium hypochlorite	$CsClO$
13.	potassium nitrite	KNO_2
14.	iron (III) sulfate	$Fe_2(SO_4)_3$
15.	mercury (II) chlorite	$Hg(ClO_2)_2$
16.	sodium nitrate	$NaNO_3$
17.	lithium perchlorate	$LiClO_4$
18.	calcium monohydrogen phosphate	$CaHPO_4$
19.	barium bisulfite	$Ba(HSO_3)_2$
20.	silver bicarbonate	$AgHCO_3$
21.	lithium dihydrogen phosphate	LiH_2PO_4
22.	chromium (III) acetate	$Cr(C_2H_3O_2)_3$
23.	tin (IV) hydroxide	$Sn(OH)_4$
24.	zinc bisulfide	$Zn(HS)_2$
25.	aluminum bisulfate	$Al(HSO_4)_3$

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

TEACHER'S GUIDE

INTRODUCTION

In this learning cycle a double displacement reaction is investigated. The balancing of equations is used during the discussion and evidence to support the conservation of atoms and mass is examined. The idea that when ions interact there is a limited number of possible interactions is developed through discussion. Filtration terminology, filtrate and residue are defined and used in the discussion and laboatory.

CONCEPT

Chemical reactions involve changes in the combinations of atoms.

EQUIPMENT

A. EXPLORATION

1. Labware
   a. Laboratory balances, 0.01 g sensitivity
   b. Beakers--50 ml, 100 ml, 250 ml, 400 ml
   c. Filter paper
   d. Funnels
   e. Ring stand with small iron ring
   f. Plastic weighing cups
   g. Glass stirring rods
   h. Wash bottles

2. Reagents
   a. Lead nitrate (60 g)
   b. Potassium iodide (60 g)

NOTE: Sodium iodide may be used to replace potassium iodide but the weights used will change. If NaI is used to replace KI, use 3.0 grams of NaI in Step 2, Page 1 of the student materials.

B. EXPANSION OF THE IDEA

1. Labware
   a. Dropper bottles--six per station
   b. 13 x 100 mm test tubes
   c. Test tube rack

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

2. Reagents

Following are five possible sets. Use one or more in your laboratory in labeled dropper bottles.

	Concentration	Formula	Weight/100 ml
Set 1	0.1M	$Ba(NO_3)_2$	2.6 g
	0.1M	$BaCl_2$	2.4 g $BaCl_2 \cdot 2H_2O$
	0.1M	Na_2CrO_4	1.6 g
	0.1M	K_2CrO_4	1.9 g
	0.2M	$NaNO_3$	1.7 g
	0.2M	KCl	1.5 g
Set 2	0.1M	Na_2SO_4	1.4 g anhydrous
	0.07M	$Al_2(SO_4)_3$	4.7 g $Al_2(SO_4)_3 \cdot 18H_2O$
	0.1M	$Sr(NO_3)_2$	2.1 g
	0.1M	$BaCl_2$	2.4 g $BaCl_2 \cdot 2H_2O$
	0.1M	$Ba(NO_3)_2$	2.6 g
	0.07M	$AlCl_3$	1.7 g $AlCl_3 \cdot 6H_2O$
Set 3	0.1M	$FeCl_3$	2.7 g $FeCl_3 \cdot 6H_2O$
	0.1M	$Co(NO_3)_2$	2.9 g $Co(NO_3)_2 \cdot 6H_2O$
	0.1M	$CoCl_2$	2.4 g $CoCl_2 \cdot 6H_2O$
	0.2M	NaOH	0.80 g
	0.2M	KOH	1.12 g
	0.2M	$NaNO_3$	1.7 g
Set 4	0.1M	$NiCl_2$	2.4 g $NiCl_2 \cdot 6H_2O$
	0.1M	$MgCl_2$	2.0 g $MgCl_2 \cdot 6H_2O$
	0.1M	Na_2SO_4	(Set 2)
	0.2M	NaOH	(Set 3)
	0.1M	$Ba(OH)_2$	(saturated soln 1.7 g)
	0.1M	$MgSO_4$	2.5 g $MgSO_4 \cdot 7H_2O$
Set 5	0.1M	$BaCl_2$	(Set 1)
	0.1M	$Sr(NO_3)_2$	(Set 2)
	0.1M	Na_2CrO_4	(Set 1)
	0.07M	$Al_2(SO_4)_3$	(Set 2)
	0.1M	K_2CrO_4	(Set 1)
	0.2M	$AgNO_3$	3.4 g

BEST COPY AVAILABLE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

EXPLORATION

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Name_____

Section_____

INTRODUCTION

This investigation is the first of several that will involve the use of substances whose formulas can be produced from combinations of the common ions that you have learned. In doing this experiment, you will need to be very careful to follow the procedure and perform the weighings correctly.

EXPLORATION

1. Weigh 3.31 g of lead nitrate—$Pb(NO_3)_2$—into a clean, dry weighing cup. Transfer this solid to your clean 100 ml beaker. Add 25 ml of distilled water to the beaker and stir with your stirring rod until the solid is completely dissolved. Rinse off the stirring rod into the beaker.

2. Weigh 3.32 g of potassium iodide—KI—into a clean, dry weighing cup. Transfer this solid to your clean 100 ml beaker. Add 25 ml of distilled water to the beaker and stir with your stirring rod until the solid is completely dissolved. Rinse off the stirring rod into the beaker.

3. Pour the $Pb(NO_3)_2$ solution into the KI solution. Rinse the 25 ml beaker into the 100 ml beaker using your wash bottle.

4. Stir the mixture in your 100 ml beaker. Describe the results. Specify any evidence for a chemical change.

5. Weigh your clean, dry 400 ml beaker (to two decimal places).

_____

_____

_____

Average _____

OR

6. Weigh a piece of filter paper to two decimal places.

_____

_____

_____

Average _____

7. Set up your funnel on a ring stand. Place your 400 ml beaker under the funnel to catch the filtrate. Fold the filter paper and place it in the funnel, and then filter the mixture from the 100 ml beaker. Rinse any remaining residue from your 100 ml beaker into the filter, using your wash bottle.

8. When all the filtrate has passed through the filter, rinse the residue with distilled water from your wash bottle. Allow this rinse water to collect in your 400 ml beaker also.

9. Remove the filter paper and residue from the funnel and carefully place them in a clean 250 ml beaker. NOTE: BE CAREFUL NOT TO LOSE ANY OF THE RESIDUE.

10. Put the 250 ml beaker with the filter paper and residue in the place designated by your teacher.

11. Label your 400 ml beaker containing the filtrate and put it in the place designated by your teacher.

12. When the filter paper and residue are dry, weigh them to two decimal places.

_____

_____

_____

Average _____

13. Compare the weight of the residue. Show your calculations.

1. through 9. Demonstrate the weighing of an exact amount of solid into a weighing cup on a balance.

Adjust the balance until it balances with the weighing cup on it. Read the weight and write it on the board.

Set the balance to the sum of the weighing cup and the desired weight of the solid. Stress the importance of accuracy during weighing.

Fold a piece of filter paper in half, then place a spoonful of reagent in the creased paper. Tap the solid into the weighing cup until balance is achieved. In order to facilitate the weighing process, it is recommended that the teacher prepare the lead nitrate and potassium iodide by grinding them with a mortar and pestle.

NOTE: If the reagent is kept clean in handling, any excess may be returned to the reagent container.

Demonstrate the method for rinsing out one beaker into another using a wash bottle.

Demonstrate the folding of filter paper and sealing it in the funnel with water from the wash bottle.

Demonstrate the set-up of the funnel in a small ring on a ring stand. Instruct the students to place their 400 ml beaker under the funnel so that splattering of the filtrate is minimized. Make sure the students understand the terms residue and filtrate.

10. Storage options for residue drying:

a. Before the residue is collected, have the students label a 250 ml beaker.

b. Direct the students to place their filter paper with the residue in a 250 ml beaker.

c. More thorough and rapid drying is achieved when the dump residues are placed in a well ventilated area, such as on the top of a ventilation blower.

The $Pb(NO_3)_2$ solution should be dried in an oven, since it decomposes at 400°C.

BEST COPY AVAILABLE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

14. When your ... at beaker ... been returned to ..., weigh the system to two decimal places.

_____

_____

_____

Average _____

15. Record observations of the beaker's content. Compute the weight of the substance contained in the beaker. Show your calculations.

16. Record your data on the first line of the table and fill in the class data on the remaining lines.

Weight of Residue	Weight of Filtrate	% Weight Change
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____

CONCEPTUAL INVENTION (THE IDEA)

THE IDEA

1. State the evidence that a chemical change has occurred in this interaction.

2. List the formulas for the ions present in the lead nitrate [$Pb(NO_3)_2$] and potassium iodide (KI) ... before you mixed them.

3. List the formulas for the ions present after you mix the lead nitrate with the potassium iodide.

4. What change has occurred with the ions to explain the chemical change you noticed when you mixed lead nitrate and potassium iodide?

5. Write a chemical equation to represent the reaction studied in this investigation.

6. Compare the weights of the reactants with the weights of the products. What does this comparison illustrate about chemical reactions?

Acceptable Answers

1. color change, precipitate formation.

2. Pb^{2+}, NO_3^-, K^+, I^-

3. Pb^{2+}, NO_3^-, K^+, I^-

4. The ions rearranged themselves to form with different partners.

5. $Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$

If NaI was substituted for KI;
$Pb(NO_3)_2 + 2NaI \rightarrow PbI_2 + 2NaNO_3$

6. The students should compare the weight of their reactants to the weight of the products—reactants weigh slightly more, the same, or slightly less than the products. In a chemical reaction, the weights stay the same.

BEST COPY AVAILABLE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

INVENTION DISCUSSION

A. Ask if the mixture of the two solutions resulted in a chemical change. Require supporting evidence from observations (color change, precipitate formation).

   1. Have students identify the reactants by name and formula.

   2. Establish that there would have to be a limited number of product possibilities, and then ask the students to supply product possibilities.

      a. Eliminate reactant recombination as a possibility.
      b. Eliminate combinations of negative ions with each other and of positive ions with each other.

   3. When products have been specified, lead the class through balancing the equation. This can be done by showing the need to conserve ions in order to arrive at charge balanced formulas.

   4. Lead a discussion to identify the products, KNO_3 and PbI_2. This will involve looking at the product weight data. Using product formulas and atomic weights, students should ultimately be able to conclude that the heavier of the two products is lead iodide. That this product is the yellow precipitate can be elicited from the class.

B. Ask the students to examine the percentage weight change column of their weight data tables. These percentage weight changes, between total reactant weight and total product weight, should approximate 0% change. The idea that weight is conserved in chemical reactions can then be stated to summarize the discusssion of this information.

C. Ask the students to state what changes had to occur in going from reactants to products in this reaction. (The ions <u>changed partners</u> or rearranged to form products.)

   1. Make the invention statement: <u>Chemical reactions involve changes in the combination of atoms.</u>

   2. Lead the class through the perspective of imagining that displacement of positive and negative ions occurs during the reaction giving rise to the name <u>double displacement</u> reaction, or <u>ion exchange</u> reaction.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

EXPANSION OF THE IDEA

Experiment

The laboratory is a nice activity for the students to predict products using laboratory evidence. Considerable preparation time is required. Select one or more of the sets of reactants listed in the Teacher's Guide.

BEST COPY AVAILABLE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

[illegible]

[illegible]

1. At your laboratory station you will find a set of six substances in labeled dropper bottles. Number each of the substances and identify the ions present.

Table II

	[illegible]	[illegible]	[illegible]
1			
2			
3			
4			
5			
6			

2. In the following procedures be careful not to contaminate the substances in the dropper bottles. (Use only the dropper provided with each bottle. Don't use the dropper to stir mixtures and don't allow the tip of the dropper to touch the mixtures.) Into a small clean test tube put 4 [illegible] drops of Substance Number 1. Add 4 [illegible] drops of Substance Number 2 and mix by agitating the test tube. Observe the mixture for any evidence of chemical reaction.

3. Making sure the test tubes you use are thoroughly clean and rinsed with distilled water, combine each of the six substances with the other five until you have tried all of the possible pairs. Enter your data in the following chart by indicating any evidence of chemical reaction you observe. Write NR (for no reaction) for any combination(s) which shows no evidence of reaction.

[illegible]

Table III

	1	2	3	4	5	6
1						
2						
3						
4						
5						
6						

[illegible]

[illegible]

[illegible]

1. For each combination of substances which results in a precipitate write a balanced chemical equation.

2. For the reactions in Question 1, identify the product(s) precipitated. Eliminate other possibilities by examining the reactions which didn't result in a precipitate. Briefly state the logic used to identify each precipitate.

[illegible]

1. and 2. Caution the students to be careful not to contaminate the substances in the dropper bottles. The test tubes must be rinsed with distilled water.

3. All of the reactions will be double displacement reactions. Some will result in a precipitate. If the student is not sure if a precipitate has formed, have him/her put the two solutions into test tubes and inspect each solution separately before combining them.

1. $FeCl_3 + 3KOH \rightarrow Fe(OH)_3 + 3KCl$
   $FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$
   $Co(NO_3)_2 + 2KOH \rightarrow Co(OH)_2 + 2KNO_3$
   $Co(NO_3)_2 + 2NaOH \rightarrow Co(OH)_2 + 2NaNO_3$
   $CoCl_2 + 2KOH \rightarrow Co(OH)_2 + 2KCl$
   $CoCl_2 + 2NaOH \rightarrow Co(OH)_2 + 2NaCl$

2. The answer to this question for a reaction like $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ would go like this: 'AgCl or $NaNO_3$ or both must be the precipitate.'

Since $NaNO_3$ was formed when x and z were combined and no precipitate resulted, then AgCl must be the precipitate.

BEST COPY AVAILABLE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Name_____

Section_____

READING 7

QUESTIONS

Predict the products and write balanced equations for each of the following reactions.

1. Magnesium metal reacts with hydrochloric acid (HCl) to produce a gas which burns explosively.

2. Lead acetate in water reacts with sodium iodide in water to produce a yellow precipitate. When the remaining liquid is boiled down, a white solid remains.

3. Silver nitrate in water reacts with potassium chromate in water to produce a red-brown precipitate.

4. When sulfur is heated in air, a suffocating, choking gas results.

188

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

5. Iron (II) sulfide reacts with sulfuric acid to produce a gas which smells like rotten eggs.

6. Lead metal reacts with sulfuric acid (H_2SO_4) to produce a white solid precipitate and a suffocating, choking gas.

7. Copper metal reacts with nitric acid (HNO_3) to produce a blue solution and a brown, choking gas.

188

8. Assign the reading: Predicting the Proper Products and its questions. Discuss the questions and significant points of the reading. The questions and acceptable answers follow.

Acceptable Answers

1. $Mg+2HCl \rightarrow H_2+MgCl_2$

2. $Pb(C_2H_3O_2)_2+2NaI \rightarrow PbI_2+2NaC_2H_3O_2$

3. $2AgNO_3+K_2CrO_4 \rightarrow Ag_2CrO_4+2KNO_3$

4. $S+O_2 \rightarrow SO_2$

5. $FeS+H_2SO_4 \rightarrow H_2S+FeSO_4$

6. $Pb+2H_2SO_4 \rightarrow PbSO_4+2SO_2+2H_2O$

7. $Cu+4HNO_3 \rightarrow Cu(NO_3)_2+2NO_2+2H_2O$

(The students should deduce from the reading that the blue solution produced is a copper solution. The identity of the brown choking gas can be found in the chart on the first page of the reading. The third product may be difficult for the students to identify. To account for all of the atoms present in the reactants, the students may suggest hydrogen gas. Since the gas produced matched the description for nitrous dioxide, hydrogen gas would be eliminated as a possibility. The teacher might suggest that water would be a good choice for the third product.)

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

TEACHER'S GUIDE

INTRODUCTION

In this investigation another type of chemical reaction is studied. During the single displacement reaction there are two reaction processes occurring. Oxidation, which is the loss or apparent loss of electrons, is one process. The other process is reduction, which is the gain or apparent gain of electrons. The Expansion of the Idea phase contains exercises in which the student balances equations using oxidation and reduction half-reactions. The term spectator ion is defined and identified in the reactions. Also, these activities provide the student an opportunity to predict the products of a reaction by using observational data. The readings, Active and Noble Metals and Activity and Chemical Reactions, will present the idea of a reactivity series to the students.

CONCEPT

Some chemical reactions involve changes in the electron structure of atoms.

EQUIPMENT

A. Exploration

1. Labware

a. 18x150 mm test tubes
b. Test tube rack
c. Dropper bottles (two for class use)

2. Reagents

a. Concentrated hydrochloric acid (one small dropper bottle full)
b. Concentrated sulfuric acid (one small dropper bottle full)
c. Copper (II) chloride solution--0.5M (200 ml)
d. Copper (II) sulfate solution--0.5M (300 ml)
e. Zinc metal pieces (60)--large zinc shot or small pieces of mossy zinc will work well

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

PRE-LAB INSTRUCTIONS

A. The teacher should go over the general precaution of adding concentrated acids to water (or solutions) and <u>not</u> the reverse procedure. If the written procedures for these experiments are followed properly, there will be no problem caused by improper addition of concentrated acid. It is, however, worthwhile for the students to be aware of the potential danger of adding water (or a solution) to a concentrated acid.

B. Reagents that do not have to be covered, such as the copper (II) chloride solutions, can be dispensed conveniently from large flasks by using a 10 ml graduated pipet with a suction bulb on it. The pipet can remain in the open flask during the work period.

INVESTIGATION 8: COMBINATIONS WITH ZINC

EXPLORATION

INVESTIGATION 8: COMBINATIONS WITH ZINC

Name_____

Section_____

INTRODUCTION

In the last investigation, the accuracy of your weight data was stressed. This investigation uses less specific amounts of substances, but your careful observations are quite important.

EXPLORATION

1. Label a clean 18 x 150 mm test tube as System #1 and dispense approximately 10 ml of copper (II) chloride solution into it.

2. Add one piece of mossy zinc metal to the $CuCl_2$ solution and observe the system carefully for 3 to 5 minutes. Record your observations.

3. Is there any evidence for a chemical change in System #1? If so, specify.

4. Slowly add 15 drops of concentrated sulfuric acid (H_2SO_4) to System #1, shake gently to mix, and record your careful observations.

5. Is there any evidence for a chemical change as a result of adding the acid to System #1? If so, specify.

186

INVESTIGATION 8: COMBINATIONS WITH ZINC

6. Set System #1 in your test tube rack and label a second clean 18 x 150 mm test tube as System #2. Dispense approximately 10 ml of copper (II) chloride solution into the System #2 test tube.

7. Add one piece of mossy zinc metal to the $CuCl_2$ solution and observe the system for 3 to 5 minutes. Record your observations.

8. Is there any evidence for a chemical change in System #2? If so, specify.

9. Slowly add 15 drops of concentrated hydrochloric acid (HCl) to System #2, shake gently to mix, and record your careful observations.

10. Is there any evidence for a chemical change as a result of adding the acid to System #2? If so, specify.

187

1. Have the students wash the test tubes and rinse with distilled water.

2. Tell the students that they may test any gases for flammability. During the discussion of Investigation 2, the flammability test for hydrogen, using a burning wooden splint, was shown to the students.

4. In using the concentrated acid, be sure to keep the dropper pointed downward. Have the test tube in a test tube clamp before adding the acid.

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

INVESTIGATION 8: COMBINATIONS WITH ZINC

11. Set System #2 in your test tube rack. After the systems have been allowed to sit for about 10 minutes, record final observations on both systems.

System #1 System #2

CONCEPTUAL INVENTION (THE IDEA)

THE IDEA

1. If you restrict yourself to the list of common ions, what products could possibly have been formed by the chemical changes you observed in System #1? Write equations for the reactions.

2. If you restrict yourself to the list of common ions, what products could possibly have been formed by the chemical changes you observed in System #2? Write equations for the reactions.

Acceptable Answers

1. (a) $ZnSO_4$, Cu, H_2

(b) Any balanced equation involving the System 1 reactants is acceptable.

NOTE: If few students suggest other possibilities for products such as S or SO_2, accept those answers. During the discussion these substances will be eliminated as possible products.

(b) Any balanced equation involving the System 2 reactants is acceptable.

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

INVENTION DISCUSSION

A. Ask the students to reiterate any evidence they had for chemical change in Step 2. With consensus that there was a chemical change, ask them to name the substances that reacted together. If water is suggested as a possibility, place a piece of zinc in a large test tube and put in some distilled water. Leave this system in view of the students throughout the discussion.

B. When copper (II) sulfate and zinc have been named as the substances that reacted, write their formulas on the board and label them as <u>reactants.</u>

C. Ask the class what the result of this chemical reaction was; that is, what was produced by the chemical reaction.

   1. This question may be difficult for the students to answer because the copper produced does not have a luster.

   2. At this time, you should produce a sample of spongy copper, dry it between paper towels, and press it in a mortar and pestle. Pass it around for the class to observe. The luster and color of copper metal should be apparent. This may be enhanced by wiping excess powdered copper out of the mortar.

   3. When the class has agreed that the substance is copper, write its formula on the board in the appropriate place.

   4. Ask the students to determine what other substance (or substances) might have been produced. Direct their attention to the formulas for the reactants.

   5. If someone suggests zinc sulfate, write its formula on the board in the appropriate place. Ask if the reaction is balanced as written. (Yes.)

D. Label the substances on the right hand side of the arrow as <u>products.</u>

E. Ask the students to describe their observations when they added the concentrated sulfuric acid to the system (Step 4). List these observations on the board.

INVESTIGATION 8: COMBINATIONS WITH ZINC

F. Ask if the sulfuric acid became involved in a chemical reaction. When consensus is reached that a chemical change has occurred, ask the students to identify the gas produced (hydrogen). Since we need to determine the equation for the reaction that occurred as a result of adding sulfuric acid, and since we know that the reaction resulted in the production of hydrogen gas, have the students suggest various reactants in the system that are available to react with the sulfuric acid. Four equation attempts will result:

$H_2SO_4+Zn \longrightarrow H_2+$

$H_2SO_4+Cu \longrightarrow H_2+$

$H_2SO_4+CuSO_4 \longrightarrow H_2+$

$H_2SO_4+ZnSO_4 \longrightarrow H_2+$

As the students suggest the second reactants, the teacher should write the reaction fragments listed above on the board. Have the students predict the remaining product or products in each reaction and then have them attempt to balance the equations. The final result will be the reaction:

$H_2SO_4+Zn \longrightarrow H_2+ZnSO_4$.

Proceed with a similar discussion for the $CuCl_2$ system (steps 6-10).

G. Focus the attention of the class on the disappearance of the color from both solutions as the reactions proceeded. Ask them to suggest reasons for this phenomenon. (This color disappearance will, of course, correspond to the reaction of all the copper compound in a given solution.)

H. If their observations indicate that the solutions were still colored when the acids were added, ask if it is possible that both reactions were occurring simultaneously in the two systems (yes). Have them defend their positions with observational evidence (further disappearance of color and gas production).

INVESTIGATION 8: COMBINATIONS WITH ZINC

I. Write on the board the following equation:

$Zn+CuSO_4 \longrightarrow ZnSO_4+Cu$

1. Have the students tell you the number of protons and electrons in the copper when it is a reactant and when it is a product.

2. Ask the students what happened to the copper and zinc in going from a reactant to a product.

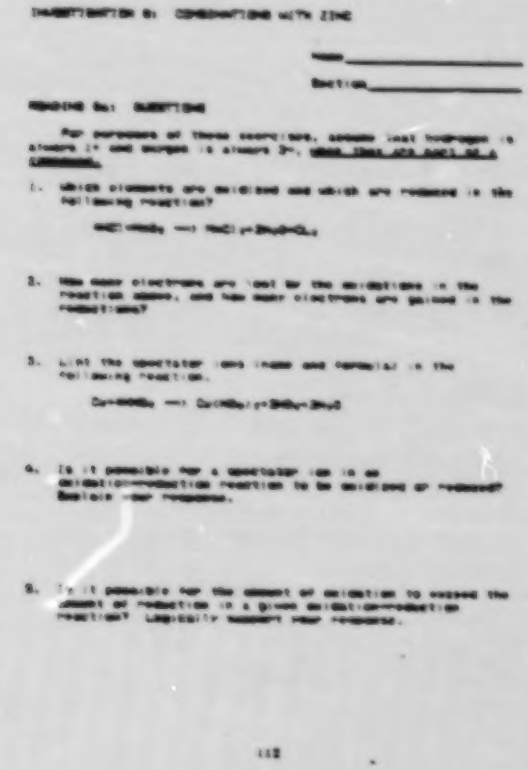
3. Once the students have stated that electrons are lost, introduce oxidation as the loss or apparent loss of electrons. Also, when the students realize electrons are gained, introduce reduction as the gain or apparent gain of electrons.

4. Ask the students what happened to the charge on the sulfate ion. Since there was no change in the sulfate ion in going from a reactant to a product, it is called a spectator ion. Introduce a <u>spectator ion</u> as an ion that is neither oxidized nor reduced by chemical reaction.

INVESTIGATION 8: COMBINATIONS WITH ZINC

EXPANSION OF THE IDEA

Assign Reading a: Active and Noble Metals, and its questions.



Name_____

Section_____

Acceptable Answers

$Mn^{+4} + 2e^- \longrightarrow Mn^{+2}$ reduced

$2Cl^- \longrightarrow Cl_2^0 + 2e^-$
oxidized

2. Two electrons were lost by each of the Cl^- ions. Two electrons were gained by the Mn^{+4} ion.

3. The hydrogen ion (H^+) and the oxide ion (O^{-2}) are spectator ions.

4. No. For the ion to be a spectator ion, the number of electrons is the same in the reactants and the products.

5. No. The electrons lost in the oxidation process are the electrons gained by the reduction process in the reaction.

After the students have completed Reading a, Active and Noble Metals, discuss the concept of oxidation and reduction half-reactions, then go through the balancing of an equation using half-reactions.

A. For the reaction $ZnS + O_2 \longrightarrow ZnO + SO_2$, write oxidation and reduction half reactions.

$S^{2-} \longrightarrow S^{+4} + 6e^-$ oxidation

$2O^0 + 4e^- \longrightarrow 2O^{2-}$ reduction

NOTE: The second half-reaction is doubled initially because there must be an even number of oxygen atoms corresponding to O_2 in the unbalanced equation.

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

B. To balance electrons in the two half-reactions:

$2[S^{2-} \longrightarrow S^{+4} + 6e^-]$

$3[2O^0 + 4e^- \longrightarrow 2O^{2-}]$

$2S^{2-} \longrightarrow 2S^{+4} + 12e^-$

$6O^0 + 12e^- \longrightarrow 6O^{2-}$

C. Adding the half-reactions and cancelling electrons:

$2S^{2-} + 6O^0 \longrightarrow 2S^{+4} + 6O^{2-}$

D. Adding in spectator ions:

$2ZnS + 3O_2 \longrightarrow 2SO_2 + 2ZnO$

Assign Reading b: Activity and Chemical Reactions and its questions.

[illegible]

Name_____

Section_____

[illegible]

Using the table in this reading, predict if the following will react.

1. [illegible]

2. [illegible]

3. [illegible]

Using the table in this reading, predict the products of the following reactions.

4. [illegible]

5. [illegible]

6. [illegible]

117

Acceptable Answers

1. Yes.
2. No.
3. No.
4. $FeBr_2 + Sn$
5. $2AgCl$
6. No reaction.

When the reading is completed, assign Problem Set 8 which has equations to balance by half-reactions.

BEST COPY AVAILABLE

INVESTIGATION 8: COMBINATIONS WITH ZINC

PROBLEM SET 8: ANSWER KEY

1. $3[S^{2-} \longrightarrow S^0 + 2e^-]$
$2[N^{5+} + 3e^- \longrightarrow N^{2+}]$

$3S^{2-} \longrightarrow 3S^0 + 6e^-$
$2N^{5+} + 6e^- \longrightarrow 2N^{2+}$

$2N^{5+} + 3S^{2-} \longrightarrow 3S^0 + 2N^{2+}$

$2HNO_3 + 3H_2S \longrightarrow 3S + 2NO + 4H_2O$

2. $2I^0 \longrightarrow 2I^{5+} + 10e^-$
$10[N^{5+} + e^- \longrightarrow N^{4+}]$

$2I^0 \longrightarrow 2I^{5+} + 10e^-$
$10N^{5+} + 10e^- \longrightarrow 10N^{4+}$

$2I^0 + 10N^{5+} \longrightarrow 2I^{5+} + 10N^{4+}$

$I_2 + 10HNO_3 \longrightarrow 2HIO_3 + 10NO_2 + 4H_2O$

3. $2I^- \longrightarrow 2I^0 + 2e^-$
$Mn^{4+} + 2e^- \longrightarrow Mn^{2+}$

$Mn^{4+} + 2I^- \longrightarrow Mn^{2+} + 2I^0$

$MnO_2 + 4HI \longrightarrow MnI_2 + I_2 + 2H_2O$

4. $4[2I^- \longrightarrow 2I^0 + 2e^-]$
$S^{6+} + 8e^- \longrightarrow S^{2-}$

$8I^- \longrightarrow 8I^0 + 8e^-$
$S^{6+} + 8e^- \longrightarrow S^{2-}$

$8I^- + S^{6+} \longrightarrow S^{2-} + 8I^0$

$8NaI + 5H_2SO_4 \longrightarrow H_2S + 4I_2 + 4Na_2SO_4 + 4H_2O$

INVESTIGATION 8: COMBINATIONS WITH ZINC

(Below: The least common multiple of 3Cl and 2Cl from the two sides is 6Cl--a good place to start.)

5. $6Cl^- \longrightarrow 6Cl^0 + 6e^-$
$2N^{5+} + 6e^- \longrightarrow 2N^0$

$2N^{5+} + 6Cl^- \longrightarrow 2N^0 + 6Cl^0$

$2NF_3 + 2AlCl_3 \longrightarrow N_2 + 3Cl_2 + 2AlF_3$

6. $3[P^{3+} \longrightarrow P^{5+} + 2e^-]$
$2[N^{5+} + 3e^- \longrightarrow N^{2+}]$

$3P^{3+} \longrightarrow 3P^{5+} + 6e^-$
$2N^{5+} + 6e^- \longrightarrow 2N^{2+}$

$2N^{5+} + 3P^{3+} \longrightarrow 2N^{2+} + 3P^{5+}$

$2HNO_3 + 3H_3PO_3 \longrightarrow 2NO + 3H_3PO_4 + H_2O$

7. $2I^0 \longrightarrow 2I^{5+} + 10e^-$
$5[Cl^+ + 2e^- \longrightarrow Cl^-]$

$2I^0 \longrightarrow 2I^{5+} + 10e^-$
$5Cl^+ + 10e^- \longrightarrow 5Cl^-$

$2I^0 + 5Cl^+ \longrightarrow 2I^{5+} + 5Cl^-$

$I_2 + 5HClO + H_2O \longrightarrow 2HIO_3 + 5HCl$

8. $4[N^{3-} \longrightarrow N^{2+} + 5e^-]$
$5[2O^0 + 4e^- \longrightarrow 2O^{2-}]$

$4N^{3-} \longrightarrow 4N^{2+} + 20e^-$
$10\ O^0 + 20e^- \longrightarrow 10O^{2-}$

$4N^{3-} + 10O^0 \longrightarrow 4N^{2+} + 10O^{2-}$

$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$

INVESTIGATION 8: COMBINATIONS WITH ZINC

9. $2[S^{2-} \longrightarrow S^0 + 2e^-]$
$S^{4+} + 4e^- \longrightarrow S^0$

$2S^- \longrightarrow 2S^0 + 4e^-$
$S^{4+} + 4e^- \longrightarrow S^0$

$2S^- + S^{4+} \longrightarrow 3S^0$

$2H_2S + H_2SO_3 \longrightarrow 3S + 3H_2O$

10. $2Cu^+ \longrightarrow 2Cu^{2+} + 2e^-$
$S^{2-} \longrightarrow S^{6+} + 8e^-$

$10[N^{5+} + e^- \longrightarrow N^{4+}]$
$2Cu^+ \longrightarrow 2Cu^{2+} + 2e^-$

$S^{2-} \longrightarrow S^{6+} + 8e^-$
$10N^{5+} + 10e^- \longrightarrow 10N^{4+}$

$2Cu^+ + S^{2-} + 10N^{5+} \longrightarrow 2Cu^{2+} + S^{6+} + 10N^{4+}$

$Cu_2S + 12HNO_3 \longrightarrow$
$Cu(NO_3)_2 + CuSO_4 + 10NO_2 + 6H_2O$
(there are two nitrates
that are spectators)

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

TEACHER'S GUIDE

INTRODUCTION

During previous investigations various types of chemical reactions have been studied. A decomposition reaction is investigated in this learning cycle. The major invention of the learning cycle is catalysis. The expansion laboratory activities develop the idea that the rate of reaction is affected by the amount of catalyst and the amount of the reactants present.

CONCEPT

Catalysts are substances that speed up a chemical reaction without undergoing a net change.

EQUIPMENT

A. Exploration

1. Labware

a. 125 ml Erlenmeyer flask
b. Plastic spoon
c. Wood splints
d. 25 ml graduated cylinders
e. Laboratory balance

2. Reagents

a. 3% hydrogen peroxide solution
b. Manganese dioxide-50g

B. Expansion of the Idea

1. Labware

a. 400 ml beaker
b. Ring stand with burette clamp and utility clamp
c. 25 x 150 mm test tube
d. Number 4, one-hole rubber stopper
e. 90° bend glass tubing
f. Rubber tubing
g. 50 ml eudiometer

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

2. Reagents

a. 3% hydrogen peroxide solution
b. Manganese dioxide

PRE-LAB INSTRUCTIONS

A. Emphasize the importance of careful cleaning of the Erlenmeyer flask used in Step 1 and of care in measuring weights throughout the laboratory.

B. Show the students how solid can be transfered to a flask with a narrow opening by rolling up a piece of paper to make a funnel.

BEST COPY AVAILABLE

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

EXPLORATION

Name_____

Section_____

INTRODUCTION

You have probably put hydrogen peroxide solution on a cut at one time or another. Do you recall the effect you observed? What caused the observed effect? How is hydrogen peroxide normally stored in a brown bottle? Your reaction of hydrogen peroxide with manganese dioxide in this investigation will help in your understanding of the answers to these questions.

EXPLORATION

1. Clean and dry your 125 ml Erlenmeyer flask and determine its weight to two decimal places.

_____

_____

_____

average _____

2. Add 1 spatulaful of manganese dioxide (MnO$_2$) to the flask and weigh the system.

_____

_____

_____

average _____

125

3. Compute the weight of MnO$_2$ in the system. Show your calculations.

4. Add 10 ml of hydrogen peroxide (H_2O_2) solution to the system. Test the gas being produced with a glowing wood splint. Record your observations.

5. When the gas production ceases, label the flask and put it in the place designated by your teacher.

6. When the flask is returned to you, weigh the system again.

_____

_____

_____

average _____

7. Determine the weight of the contents of the flask. Show your calculations.

8. Record your observations of the contents of the flask.

125

4. Tell the students that they may test any gases produced for flammability.

Caution the students not to allow ashes from the wood splint to get into the Erlenmeyer flask since these ashes will affect weight readings. Suggest that the flammability tests be performed while holding the flask at an angle. This procedure will reduce the hazard of getting fingers burned from the flame on the wood splint.

5. The flasks should be oven-dried to preclude the decomposition of manganese dioxide due to strong heating. Decomposition occurs at 535°C.

BEST COPY AVAILABLE

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

CONCEPTUAL INVENTION (THE IDEA)

[redacted]

Acceptable Answers

1. The percentage weight change should be near zero.

2. The contents of the flask are black and similar in appearance to the MnO_2. Since the weight of the contents of the flask have changed less than 1%, the contents of the flask must be MnO_2.

3. (a) Any reasonable response, such as the oxygen came from the H_2O_2 or the MnO_2. The statement should indicate that the oxygen came from the system.

(b) Accept any balanced equation which has oxygen as a product and accounts for the reactants.

INVENTION DISCUSSION

A. Ask if a chemical reaction (or change) occurred in the H_2O_2/MnO_2 system. (Chemical change occurred, as evidenced by the production of gas. Occasionally, though not often, heat production will be cited as evidence.)

B. Ask the students to identify the gas. (The gas is oxygen and most students have no trouble deciding this. There is sometimes a small popping noise, when the splint begins burning with a flame, and this noise causes some students to suggest erroneously that hydrogen is the gas being produced. Discussion of previous tests for hydrogen can help clear up the problems encountered here.)

1. Have the students specify the source of the oxygen (O_2) gas.

2. Write down proposed sets of reactants and products for as many equations as are proposed.

BEST COPY AVAILABLE

3. Balance, or attempt to balance, each of the relationships stated. Some relationships fall from discussion because they cannot be balanced. Others can be balanced, but further analysis will show that they must be discarded. Students may suggest relationships that can be balanced, for example:

   $H_2O_2+MnO_2 \longrightarrow H_2O+MnO+O_2$

   $2H_2O_2+MnO_2 \longrightarrow 2H_2O+Mn+2O_2$.

   Either of these <u>could be</u> correct from the standpoint of oxygen gas production and a solid remaining in the flask, but neither of them will work when analyzed with respect to the weight data on the systems. It is at this time that the weight data, especially the percentage weight change, should be analyzed. When the students have agreed that the weight change is zero (for all practical purposes), have the students dictate the reactants and products. They may wish to continue putting MnO_2 on both sides of the equation. This procedure can be corrected in time.

C. Since MnO_2 appears on both sides of the equation and since all evidence points to the fact that the same amount of MnO_2 is present before and after the reaction, ask the students if H_2O_2 would have reacted without the MnO_2. (This discussion usually involves talking about students' experience with bottles of medicinal hydrogen peroxide solution. Usually, someone knows of a time that a bottle has been left open and has gone <u>flat.</u>)

   1. When there is consensus that the reaction described by the equation will occur without the manganese dioxide, ask what effect the addition of MnO_2 has on the reaction. (The MnO_2 speeds up the reaction.)

   2. Point out (invent) that MnO_2 acts as a catalyst in this reaction. <u>Catalysts are substances that increase the rate of a chemical reaction without undergoing a net change in the overall reaction.</u>

   3. Tell the students that catalysts are not written with the reactants or products because they undergo no net change, but their formulas are often placed over the arrow in the equation to show that they are present.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

D. With the equation for the reaction on the board, ask the students what this type of reaction could be called. Invent <u>decomposition</u> reaction.

$$2H_2O_2 \xrightarrow{MnO_2} 2H_2O + O_2$$

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

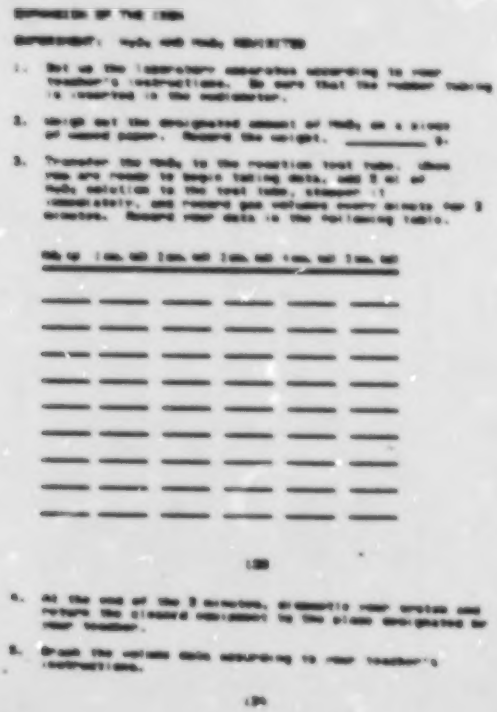
EXPANSION OF THE IDEA: EXPERIMENT

A. Talk through a mock-up of the apparatus and instruct the students to agitate the mixture in the test tube gently throughout the procedure. Stress precise laboratory technique.

H_2O_2+MnO_2

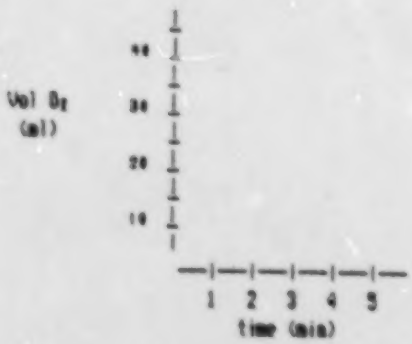
B. Assign weights of MnO_2 beginning with 0.1 g and increasing in 0.1 g increments, making sure to include values such as 1.0 g, 2.0 g, 3.0 g, etc. The weights assigned will, of course, be influenced by the number of laboratory groups in a given class.

C. The MnO_2 should be put in the test tube <u>before</u> the H_2O_2 is poured into the system. Stress to the students the importance of having the rubber tubing inserted in the eudiometer prior to pouring in the H_2O_2. The stopper should be put on the test tube as soon as the H_2O_2 solution is added. Timing begins with the addition of the solution.



5. Put data from all groups on the board, and after the students have copied all the data, assign graphing of each set. (All graphing can be done on the same sheet of graph paper. Each set of data can be labeled according to MnO_2 weight used.

Plot data on graph on overhead.



BEST COPY AVAILABLE

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

SAMPLE DATA FROM EXPANSION OF THE IDEA LABORATORY

Table 1

MnO_2(g)	1 min.(ml)	2 min.(ml)	3 min.(ml)	4 min.(ml)	5 min.(ml)
0.1	1.4	2.6	3.5	4.3	5.1
0.2	2.8	4.6	5.8	7.1	7.6
0.3	4.9	7.2	8.8	10.2	11.3
0.5	5.9	8.5	10.4	11.8	13.3
1.0	8.5	12.4	14.4	16.1	17.1
1.5	11.8	14.8	17.5	19.8	21.8
2.0	12.8	17.8	20.2	22.7	24.8
2.5	13.6	19.8	22.1	24.7	27.3
3.0	16.2	21.8	25.1	28.2	30.4

Below is a graph of data in Table 1. It shows the relationship between volume of evolved oxygen vs. time for varying amounts of MnO_2.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

[illegible]

QUESTIONS

1. [illegible]

2. Is there a relationship between the amount of catalyst and the rate of reaction? If so, state and explain the relationship. If not, explain why there is no relationship.

Acceptable Answers

1. Graphs are not always ideal; however, the general patterns will exist.

There is a rapid gassing rate during the first minute with a slower rate thereafter. The initial rate is greater for larger amounts of MnO_2.

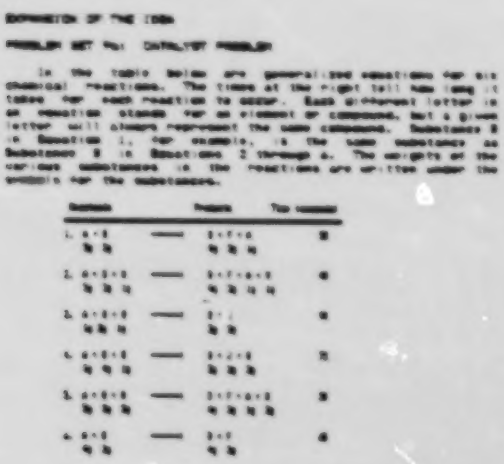
2. Discussion should proceed to answer the questions asked of the students at the end of the procedure. Elicit explanations for the differences in slopes from 0 to 1 minute compared to the remaining time. (The reason relates to the decreasing concentration of H_2O_2 as the reaction goes along.)

The result of this discussion should be the conclusion that the amount of catalyst used will affect the rate of reaction. Possible <u>thumbnail</u> mechanisms for catalysis can be explored, if the teacher so desires.

BEST COPY AVAILABLE

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Assign Problem Set a, Catalyst Problem. The discussion of this problem, after the students have worked through it, will help to reinforce the ideas concerning catalysts. The problems and acceptable answers follow.



1. Reaction #2: A+B+H —) D+F+G+H
   Reaction #5: A+B+H —) D+F+G+H

A catalyst, by definition, speeds up a reaction and is not used up in the overall reaction. Therefore, the catalyst and its weight would have to appear both as a reactant and as a product.

2. We would assume Reactions #2 and #5 occur in less time due to the presence of the catalyst.

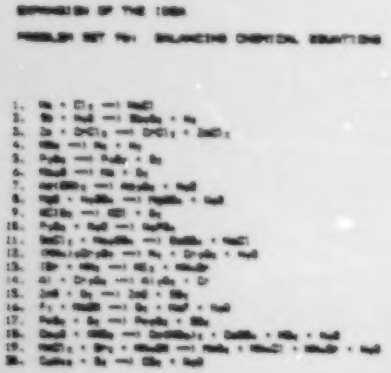
In comparing Reaction #1 and Reaction #6 the same reactants and products are present. In Reaction #1, 1g of Reactant A is in excess. Reaction #6 might have taken longer due to the fact that there were fewer Reactant A particles with which Reactant B could collide.

BEST COPY AVAILABLE

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Assign Reading 9, Enzymes: The Catalysts of Living Systems at this time. The teacher should hold a class discussion of the salient points of the reading.

Assign Problem Sets b and c, Balancing Chemical Equations now. Discussion of all three problem sets will be valuable.



BEST COPY AVAILABLE

INVESTIGATION 10: MASS RELATIONSHIPS

TEACHER'S GUIDE

INTRODUCTION

The work necessary to lead up to a study of the mole concept has been done through experimentation and through the development of the model for the atom that allows the writing of simple compound formulas by assuming the combination of oppositely-charged ions. The mole concept will be invented as a result of discussion of laboratory data, and then the students will be exposed to written exercises which will enhance their understanding of the concept.

CONCEPT

Chemical substances react in definite molar weight proportions.

EQUIPMENT

A. Exploration

1. Labware

   a. Funnels
   b. Beakers-150 ml and 250 ml
   c. Graduated cylinders-10 ml and 25 ml
   d. Glass stirring rod
   e. Filter paper to fit funnel
   f. Ring stand with small ring to hold funnel
   g. 13 x 100 mm test tubes--2
   h. Test tube rack

2. Reagents

   a. Sodium chloride solution (100 g NaCl in one liter of solution)--400 ml
   b. Silver nitrate solution (100 g $AgNO_3$ in one liter of solution)--400 ml. This solution should be kept in a brown bottle.

PRE-LAB INSTRUCTIONS

The high cost of silver nitrate solid indicates that caution should be used in its preparation and use. Tell the students the silver nitrate will stain their clothing and skin. Have them thoroughly clean up all drips and spills.

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

INTRODUCTION

This investigation involves a double displacement reaction in which you will determine one of the reactant weights, given information about the other reactant. It is important that you make CHEMICAL predictions about the reactant weights. As an aid to this process, you may refer to the balanced equation for the reaction (Step 2) and to the experiment you did in Investigation 7. Care in handling the chemicals and care in weighing the products will enhance your results.

EXPLORATION

1. Obtain about 1 ml of AgNO$_3$ solution and 1 ml NaCl solution in separate, clean test tubes. Combine the two solutions and record your observations.

2. Write a balanced equation for the chemical reaction observed.

3. Identify any precipitates formed in this reaction. Provide evidence to support your identification.

4. You will be using 30.0 ml of AgNO$_3$ solution labeled 100 g AgNO$_3$ per liter. How many grams of AgNO$_3$ will be in your 30.0 ml sample? Show your work.

5. Determine the number of grams of NaCl necessary to react with all of the AgNO$_3$ available in the reaction. Show your work below and check with your teacher before proceeding.

6. Determine the number of ml of NaCl solution corresponding to the weight of NaCl in Step 5. You will use a solution labeled 100 grams NaCl per liter. Show your calculations below.

7. Summarize the amounts of AgNO$_3$ and NaCl determined in the table below.

	weight	volume
AgNO$_3$		30.0 ml
NaCl		

8. Clean your 100 ml and 250 ml beakers with soap and water, rinse them with tap water, and then rinse the beakers with distilled water.

9. Dry and weigh two 250 ml beakers to two decimal places.

_____

_____

_____

average _____

2. and 3. It will take a whole class period (55 minutes) for the students to work through Steps 1 through 7. The teacher needs to work with the students in a non-directive manner. They will want the answer provided, but they should work it out themselves.

If they consider the equation for the reaction, a reasonable hypothesis can be made. The important point is for the students to think about the reaction and for them to make a reasonable hypothesis even if it is wrong.

4. The students will need to complete Steps 4 through 7 prior to going back to the laboratory. The students will need encouragement to think through these questions on their own. No clue should be given as to the CORRECT way to calculate the weight of reactants needed.

7. Since the reactant weights (hypothesized by the students) in this exercise will be used by them in this experiment, the teacher should check the predictions prior to the students beginning Steps 8 and following.

BEST COPY AVAILABLE

10. Call the attention of the students to the CAUTION about the silver nitrate solution. Instruct the students, by example, on the proper method for disposing these reagents.

11. After the solutions (sodium chloride and silver nitrate) have been poured into the beaker, the students are directed to rinse the graduated cylinders into the reaction beaker with small amounts of distilled water. This is to insure that all of each reactant is transferred to the beaker. Even so, the graduated cylinders should be cleaned after use.

12. The filtering procedure will require less time if the precipitate is allowed to settle. The filtrate can be poured through the filter and the major part of the solid can be transferred by pouring, scraping with the glass rod, and rinsing with distilled water.

The residue (silver chloride) should be rinsed thoroughly—at least two times—with distilled water.

NOTE: It would be wise to remind the students to weigh the filter paper before filtering the residue. They may wish to weigh the paper while the precipitate is settling in the beaker.

After the solid has been rinsed, the paper and residue should be removed carefully from the funnel and set carefully in a clean, labeled beaker. The beaker containing the residue and filter paper can be dried over warm forced air. The beakers containing the filtrates should be dried overnight in a low temperature oven.

BEST COPY AVAILABLE

INVESTIGATION 10: MASS RELATIONSHIPS

CONCEPTUAL INVENTION (THE IDEA)

[illegible]

1. [illegible]

2. [illegible]

3. [illegible]

Acceptable Answers

1. The amount of AgCl produced was close to the same regardless of the amount of NaCl used.

2. The weight of the reactants equals (or is close) the weight of the products for most groups. These data support the Law of Conservation of Weight.

3. Using the data table, the students should be able to determine that the systems using 2 g of NaCl (or more) had excess NaCl.

INVENTION DISCUSSION

A. The teacher should make headings on the board that correspond to the headings of the table in the student laboratory handout.

B. Initial discussion should center around the columns containing data on the predicted and actual weights of silver chloride. Many students will probably have hypothesized that 2.0 grams of silver chloride would form. In these cases the actual weight data will be significantly less than the predicted weights. Some students will have discerned that it is necessary to account for the different formula weights for the reactants. The predicted and actual weights for these should be the same within balance uncertainty. In reality, however, the actual weights often come out low. If these weights are not the same, the teacher will (at the appropriate time in the discussion) want to elicit the reasons for any discrepancies.

NOTE: The remaining guidelines for this discussion will assume that no student has made predictions based on comparisons of formula weights. If some students predict the product weight correctly, the discussion should be handled so that these data can be used to enhance the

BEST COPY AVAILABLE

development of the concept. Regardless of the nature of the data display, the essence of this set of guidelines remains valid for properly developing the concept. These points should, therefore, be incorporated into the discussion.

C. The data should be grouped according to sets of common reactant weights, predicted product weights, and average actual product weights. This operation will allow the teacher to summarize the class data, leave the summary on the board, and still have sufficient board space to use during the discussion.

D. The teacher should have students explain their reasons for predicting a particular weight of silver chloride. These reasons can be summarized on the board.

E. Using the list of reasons as a focus for discussion will allow the teacher to bring the class to the point of understanding that the assumption of a 2 gram weight of silver chloride implies that each of the reactants and products has the same weight.

1. Ask the class how the molecule or <u>unit</u> weight of each reactant and product would be determined--an operation that they should be able to perform. When the weights have been determined, ask what the units of these weights would be (that is, atomic mass units).

2. Propose that the same weight relationships be applied to the gram weights of reactants and products in this experiment. For example:

$$\frac{\text{wt. of } AgNO_3 \text{ (table)}}{\text{wt. of NaCl (table)}} = \frac{\text{wt. of } AgNO_3 \text{ (grams)}}{\text{wt. of NaCl (grams)}}$$

3. Perform the similar mathematical operations to determine the amount of silver chloride that would be produced.

4. A final application of these operations to the data can be made by calculating the amount of sodium nitrate that should form. Compare these results to the experimental values obtained by the students.

F. The students' results should indicate conservation of weight in the reaction, regardless of the NaCl weight chosen at the beginning. The following experiment will be useful in determining whether all the $AgNO_3$ and/or NaCl has been reacted.

1. Have the students clean two 18 x 150 mm test tubes and rinse them with distilled water.

2. A small sample of <u>filtrate</u> residue should be dissolved in each test tube by adding distilled water and stirring with a clean stirring rod.

3. Approximately two milliliters of NaCl solution should be added to one of the test tubes, and approximately two milliliters of $AgNO_3$ solution should be added to the other test tube. Stress to the students the importance of knowing which solution was added to each test tube.

4. Have the students shake the test tubes gently to mix (without stoppering the test tubes).

5. Results of these experiments will depend on the original weight of NaCl chosen as follows:

   a. 2.0 g of NaCl: When $AgNO_3$ is added, a large amount of precipitate will form. When NaCl is added, no precipitate will form.

   b. 0.69 g of NaCl: When $AgNO_3$ is added, very little (if any) precipitate will form. When NaCl is added, no precipitate will form.

NOTE: The relatively small amount of precipitate with the addition of $AgNO_3$, in the second case, can be attributed to excesses caused by lack of precision in measuring the NaCl solution volume during the initial experimental procedure.

6. Lead the class in interpreting the results of this experiment (that is, all the $AgNO_3$ was reacted in both procedures, but there is excess NaCl when 2.0 g of NaCl are used in the experiment).

G. Ask the class to predict the weight of silver chloride that would be produced by the reaction of 170 g of silver nitrate with sufficient sodium chloride. When they have properly answered this question, ask them how much sodium chloride would be sufficient; then, ask how much sodium nitrate would be produced. Work through these calculations, as necessary.

H. When the class understands that the weights suggested are table weights of the substances with units of grams, invent the term mole to describe this quantity.

I. Refer the students to the equation for the reaction of silver nitrate with sodium chloride.

1. Ask them to read the equation, assuming that one mole of silver nitrate reacts.

2. Have the students read through the equation, given that fractions or multiples of moles of silver nitrate are reacted (for example, one-fourth, one-tenth, two, etc.).

3. Choose another equation (one that does not contain all unit coefficients; for example, the reaction of nitrogen and hydrogen to form ammonia) and go through the same equation reading process as above.

J. Ask the class to compare the number of molecules in a mole of silver nitrate, sodium chloride, silver chloride, and sodium nitrate. Allow time for students to think about this and be sure that all students get the idea before pursuing it further. (Since we have balanced equations in terms of numbers of molecules, this idea can now be understood.)

K. When the class understands that a mole of any of these substances contains the same number of particles, ask them why the weights are different (the weights of different molecules are not the same).

NOTE: This type of questioning should continue until the teacher feels that the students have a reasonable first presentation understanding of these ideas. The important understanding, at this point, is simply that the number of particles in a mole of a substance is independent of the substance under discussion.

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

A. Assign Reading 11: <u>You Won't Find Any Avogadro in a Guaca-Mole Salad.</u> It provides information that will serve to reinforce the mole concept.

B. The next exercise is a series of written exercises designed to help the students understand the concept of the mole. Introduce the first set of mole problems, Problem Set a: Molar Conversions. The teacher should work a few examples and then the problem set should be assigned as homework.

In addition to the first exercise sheet, the students should do the following sets of exercises:

Problem Set b: Percent Composition and Simplest Formulas

Problem Set c: Simplest Formulas

Problem Set d: Mole in Chemical Reactions (Computation of the number of moles [and/or grams] produced or reacted according to some reaction information)

Problem Set e: Reactions with One Reactant in Excess (Problems involving excess quantities of reactants.

The teacher should give introductory examples for the various problem types. These exericises may be done in class and/or as homework assignments, as the level of understanding dictates.

NOTE: Solution concentration will be taken up as part of INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES.

INVESTIGATION 10: MASS RELATIONSHIPS

PROBLEM SET a: MOLAR CONVERSIONS

DIRECTIONS: Show all work. Use units. Answer to the nearest tenth.

1. How many grams are in 8.125 moles of Fe?

$$\frac{56\ g}{mole} = \frac{x}{8.125\ mole} \qquad x = \frac{(56\ g)(8.125\ mole)}{mole}$$

$x = 7\ g$

2. 98 grams represents how many moles of calcium?

$$\frac{40\ g}{mole} = \frac{98\ g}{x} \qquad x = \frac{98\ g}{40\ g}$$

$x = 2.25$ mole

3. How many grams are in 1.2 moles of gallium?

$$\frac{x}{1.2\ mole} = \frac{70\ g}{mole}$$

$x = 84\ g$

4. What is the weight of 2 moles of magnesium oxide?

$MgO = 24\ g + 16\ g$
$= 40\ g$

$$\frac{40\ g}{mole} = \frac{x\ g}{2\ mole} \qquad x = 80\ g$$

5. 149 grams is equivalent to how many moles of calcium(II)chloride?

$$\frac{129\ g}{mole} = \frac{149\ g}{x\ mole}$$

$CaCl_2 = 39\ g + 2(35.5g)$
$= 39\ g + 71\ g$
$= 138\ g$

$$x = \frac{149\ g}{138\ g} = 1.3 \text{ moles}$$

6. How many grams are in 3.8 moles of magnesium phosphate?

$$\frac{262\ g}{1\ mole} = \frac{x}{3.8\ mole}$$

$Mg_3(PO_4)_2$ = 3(24 g)+2(31 g)+8(16 g)
= 72 g + 62 g + 128 g
= 262 g

x = 995.6 g

7. How many grams are in 6.5 moles of sodium acetate?

$$\frac{82\ g}{1\ mole} = \frac{x}{6.5\ mole}$$

$NaC_2H_3O_2$=23 g+2(12 g)+3(1 g)+2(16 g)
= 23 g+24 g+3 g+32 g
= 82 g

x = 533 g

8. 324.7 grams of magnesium chlorate is how many moles?

$$\frac{191\ g}{mole} = \frac{324.7\ g}{x\ mole}$$

$Mg(ClO_3)_2$=24 g+2(35.5 g)+6(16 g)
= 24 g+71 g+96 g
= 191 g

x = 1.7 mole

9. How many grams are in 8.85 moles of aluminum hydroxide?

$$\frac{78\ g}{mole} = \frac{x}{8.85\ mole}$$

$Al(OH)_3$=27 g+3(16 g)+3(1 g)
= 27 g+48 g+3 g
= 78 g

x = 3.9 g

10. How much do 5.2 moles of strontium thiocyanate weigh?

$$\frac{284\ g}{mole} = \frac{x}{5.2\ mole}$$

$Sr(SCN)_2$=88 g+2(32 g)+2(12 g)+2(14 g)
= 88 g+64 g+24 g+28 g
= 284 g

x = 1868.8 g

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

PROBLEM SET b: PERCENT COMPOSITION & SIMPLEST FORMULAS

1. Determine the percentage by weight of potassium and iodine in potassium iodide.

   KI = 39 g+127 g=166 g/mole

   $$\%K = \frac{39\ g}{166\ g} = 0.2349 = 23.49\%$$

   $$\%I = \frac{127\ g}{166\ g} = 0.765 = 76.5\%$$

2. What is the percentage by weight of magnesium and oxygen in magnesium oxide?

   MgO = 24 g+16 g = 40 g

   $$\%Mg = \frac{24\ g}{40\ g} = 0.6 = 60\%$$

   %O = 16 g/40 g = 0.4 = 40%

3. Find the percentage by weight of each element in strontium chloride.

   $SrCl_2$ = 88 g+ 2(35.5)g = 159 g/mole

   %Sr = 88 g/159 g = 55.34%

   %Cl = 71 g/159 g = 44.65%

4. What is the percentage by weight of copper, carbon, and oxygen in copper (II) oxalate?

   CuC_2O_4 = 64 g+2(12 g)+4(16 g)
   = 64 g+24 g+64 g
   = 152 g

   %Cu = 64 g/152 g = 42.1%
   %C = 24 g/152 g = 15.8%
   %O = 64 g/152 g = 42.1%

5. Determine the percentage by weight of the elements in ammonium hydroxide.

   NH_4OH = 14 g+4(1 g)+16 g+1 g
   = 14 g+4 g+16 g+1 g
   = 35 g

   %N = 14 g/35 g = 40%
   %H = 5 g/35 g = 14.2%
   %O = 16 g/35 g = 45.7%

6. Find the percentage by weight of chromium, hydrogen and oxygen in chromium (III) acetate.

$Cr(C_2H_3O_2)_3$

= 52 g+6(12 g)+9(1 g)+6(16 g)
= 52 g+72 g+9 g+96 g
= 229 g

%Cr = 52 g/229 g = 22.7%
%C = 72 g/229 g = 31.4%
%H = 9 g/229 g = 3.9%
%O = 96 g/229 g = 41.9%

7. A compound contains 36% nickel and 64% chlorine by weight. What is the simplest formula for this substance?

Ni = 59g/mole Cl = 35.5g/mole 36gNi + 64gCl = 100g Ni_xCl_x

$$Ni: \frac{59g}{mole} = \frac{36g}{x\ mole} \qquad Cl: \frac{35.5g}{mole} = \frac{64g}{x\ mole}$$

x = 0.61 mole x = 1.8 mole

$$\frac{0.61\ mole}{0.61\ mole} = 1 \qquad \frac{1.8\ mole}{0.61\ mole} = 3 \qquad Ni_xCl_x = NiCl_3$$

8. Find the simplest formula for boron carbide, a compound which is made of 78.4% boron and 21.4% carbon.

B = 11 g/mole C = 12 g/mole 78.4g B + 21.4g C → B_xC_x

$$\frac{11g}{mole} = \frac{78.4g}{x} \qquad \frac{12g}{mole} = \frac{21.4g}{x}$$

x = 7.13 mole x = 1.78 mole

$$\frac{7.13\ mole}{1.78\ mole} = 4.02 \qquad \frac{1.78\ mole}{1.78\ mole} = 1 \qquad \therefore B_4C$$

9. Calculate the formula for scandium nitrate which contains 19.5% Sc, 18.2% N, and 62.3% O.

19.5 g Sc + 18.2 g N + 62.3 g O → $Sc_x(NO_3)_z$

At. wts.					
Sc = 45g	45g / mole	=	19.5g / x mole	14g / mole	= 18.2g / x mole
N = 14g					
O = 16g					

x = 0.43 mole x = 1.3 mole

$$\frac{62.3g}{x\ mole} = \frac{16g}{mole} \qquad \frac{1.3\ mole}{0.43\ mole} = 3.02 \qquad \frac{3.89\ mole}{0.43\ mole} = 9.04$$

x = 3.89 mole

$$\frac{0.43\ mole}{0.43\ mole} = 1$$

∴ $Sc(NO_3)_3$

18. A compound contains 78.6% boron and 21.4% hydrogen. What is its simplest formula?

B = 11g/mole H = 1g/mole 78.6g B + 21.4g H → B_xH_z

78.6g / x	= 11g / 1 mole	21.4g / x	= 1g / 1 mole	7.15 mole / 7.15 mole = 1

x = 7.15 mole x = 21.4 mole

$$\frac{21.4\ mole}{7.15\ mole} = 2.99 = 3$$

∴ B_3H

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

PROBLEM SET C: SIMPLEST FORMULAS

1. Acetic acid has the formula CH_3COOH. By weight what percentage of the compound is carbon? What percentage is oxygen? Hydrogen?

C = 12g
H = 1g
O = 16g

$2C + 2H + 4O = 2(12g) + 2(16g) + 4(1g) = 60g$

$$C = \frac{24g}{60g} = 0.4 \times 100 = 40\%$$

$$H = \frac{4g}{60g} = 0.066 \times 100 = 6.6\%$$

$$O = \frac{32g}{60g} = 0.533 \times 100 = 53.3\%$$

2. Bromerite contains 43.5% gold and the remainder is tellurium. What is its formula?

Au = 197g Te = 127.6g 43.5g Au + 56.5g Te = Au_xTe_y

$$\frac{43.5g}{x} = \frac{197g}{1\ mole} \qquad \frac{56.5g}{x} = \frac{128g}{1\ mole} \qquad \frac{0.22\ mole}{0.22\ mole} = 1$$

$x = 0.22$ mole Au $x = 0.44$ mole Te $\frac{0.44\ mole}{0.22\ mole} = 2$ . . $AuTe_2$

3. Hydroxylamine is 9% hydrogen, 48.5% oxygen and the remainder is nitrogen. What is its simplest formula?

H = 1g
O = 16g
N = 14g

9g H + 48.5g O + 42.5g N —) $H_9O_3N_3$

$$\frac{9g}{x} = \frac{1g}{1\ mole} \qquad \frac{48.5g}{x} = \frac{16g}{1\ mole} \qquad \frac{42.5g}{x} = \frac{14g}{1\ mole}$$

$x = 9$ mole $x = 3.03$ mole $x = 3.03$ mole

$\frac{9\ mole}{3.03\ mole} = 2.97 = 3$ $\frac{3.03\ mole}{3.03\ mole} = 1$ $\frac{3.03\ mole}{3.03\ mole} = 1$ . . HNO

4. A sample of bromodichloromethane contains 3 grams of carbon, 24 grams of bromine, 0.25 grams of hydrogen and 17.7 grams of chlorine. What is its simplest formula?

3g C + 24g Br + 0.25g H + 17.7g Cl = $HC_xBr_yCl_z$

C = 12g
Br = 80g
H = 1g
Cl = 35.5g

$\frac{3g}{x} = \frac{12g}{1\ mole}$ $\quad$ $\frac{24g}{x} = \frac{80g}{1\ mole}$

$x = 0.25$ mole $\quad$ $x = 0.25$ mole

$\frac{0.25g}{x} = \frac{1g}{1\ mole}$ $\quad$ $\frac{17.7g}{x} = \frac{35g}{1\ mole}$

$x = 0.25$ mole $\quad$ $x = 0.5$ mole

$\frac{0.25\ mole}{0.25\ mole} = 1$ $\quad$ $\frac{0.25\ mole}{0.25\ mole} = 1$ $\quad$ $\frac{0.25\ mole}{0.25\ mole} = 1$ $\quad$ $\frac{0.5\ mole}{0.25\ mole} = 2$

$HCBrCl_2$

5. Write the simplest formula for an ancient chrysoberyl composed of 2 grams of beryllium, 12 grams of aluminum and 14 grams of oxygen ($Be_xAl_yO_z$).

2g + 12g + 14g = $Be_xAl_yO_z$

Be = 9g
Al = 27g
O = 16g

$\frac{2g}{x} = \frac{9g}{1\ mole}$ $\quad$ $\frac{27g}{1\ mole} = \frac{12g}{x}$ $\quad$ $\frac{14g}{x} = \frac{16g}{1\ mole}$

$x = 0.22$ mole $\quad$ $x = 0.44$ mole $\quad$ $x = 0.875$ mole

$\frac{0.22\ mole}{0.22\ mole} = 1$ $\quad$ $\frac{0.44\ mole}{0.22\ mole} = 2$

$\frac{0.875\ mole}{0.22\ mole} = 3.97 = 4$ $\quad$ $BeAl_2O_4$

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

PROBLEM SET d: MOLE IN CHEMICAL REACTIONS

1. A mixture of hydrogen gas and fluorine gas would result in a violent reaction through which hydrogen fluoride is formed. (a) Write a balanced equation for the reaction. (b) How many moles of hydrogen fluoride will form from one mole of hydrogen gas with sufficient fluorine? (c) How many moles of fluorine are needed to produce 16 moles of hydrogen fluoride?

a) $H_2 + F_2 \longrightarrow 2HF$

b) $\frac{1 \text{ mole } H_2}{2 \text{ mole } HF} = \frac{1 \text{ mole } H_2}{x HF}$ $x = 2$ mole HF

c) $\frac{1 \text{ mole } F_2}{2 \text{ mole } HF} = \frac{x \text{ mole } F_2}{16 \text{ mole } HF}$ $x = 8$ moles

2. Crystals of potassium chlorate will decompose at high temperatures to produce oxygen gas and potassium chloride. (a) Write a balanced equation for this decomposition reaction. (b) How many moles of potassium chloride will be given by 0.66 moles of potassium chlorate? (c) How many grams of oxygen will be produced by 612.5g of potassium chlorate?

a) $2KClO_3 \xrightarrow{\Delta} 3O_2(g) + 2KCl$

b) $\frac{2 \text{ mole } KClO_3}{2 \text{ mole } KCl} = \frac{0.66 \text{ mole } KClO_3}{x \text{ mole } KCl}$ $x = 0.66$ mole

c) $KClO_3 = 39g + 35.5g + 3(16g) = 122.5$ g/mole

$\frac{122.5g}{1 \text{ mole}} = \frac{612.5g}{x}$ $\frac{2 \text{ mole } KClO_3}{3 \text{ mole } O_2} = \frac{5 \text{ mole } KClO_3}{x O_2}$

$x = 4.99$ mole $x = 7.5$ mole O_2

$\frac{32g}{1 \text{ mole } O_2} = \frac{x}{7.5 \text{ mole}}$

$x = 240$ g O_2

3. The decomposition of calcium carbonate results in the formation of calcium oxide and carbon dioxide. (a) Write a balanced equation for this decomposition reaction. (b) How many moles of calcium oxide will be produced by 30g of calcium carbonate? (c) How many moles of carbon dioxide will be formed by 212.0g of calcium carbonate?

a) $CaCO_3 \rightarrow CaO + CO_2$

b) $40g + 12g + 3(16g) = 100g$

$$\frac{30g}{1\ mole} = \frac{100g}{x} \qquad \frac{1\ mole\ CaCO_3}{1\ mole\ CaO} = \frac{0.3\ mole\ CaCO_3}{x}$$

$x = 0.3\ mole$ $\qquad x = 0.3\ mole\ CaO$

c) $$\frac{212.0g}{x\ mole} = \frac{100g}{1\ mole} \qquad \frac{1\ mole\ CaCO_3}{1\ mole\ CO_2} = \frac{2.12\ mole\ CaCO_3}{x\ mole\ CO_2}$$

$x = 2.12\ mole$ $\qquad x = 2.12\ mole\ CO_2$

4. When magnesium reacts with dilute sulfuric acid, the products formed are magnesium sulfate and hydrogen. (a) Write a balanced equation for this reaction. (b) What kind of reaction is this? (c) How many grams of magnesium sulfate will be produced by 72 grams of magnesium? (d) How many grams of hydrogen will be produced by 4.8 grams of magnesium?

a) $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2(g)$

b) Single Displacement

c) $Mg = 24g$ $\qquad MgSO_4 = 24g + 32g + 4(16g) = 120g$

$$\frac{24g}{mole} = \frac{72g}{x\ mole} \qquad \frac{1\ mole\ Mg}{1\ mole\ MgSO_4} = \frac{3\ mole\ Mg}{x\ mole\ MgSO_4}$$

$x = 3\ mole\ Mg$ $\qquad x = 3\ mole\ MgSO_4$

$$\frac{120g\ MgSO_4}{1\ mole} = \frac{x\ g\ MgSO_4}{3\ mole\ MgSO_4}$$

$x = 360g\ MgSO_4$

d) $$\frac{24g}{mole} = \frac{4.8g}{x} \qquad \frac{1\ mole\ Mg}{1\ mole\ H_2} = \frac{0.2\ mole\ Mg}{x\ mole\ H_2} \qquad \frac{2g}{1\ mole\ H_2} = \frac{x}{0.2\ mole}$$

$x = 0.2\ mole$ $\qquad x = 0.2\ mole\ H_2$ $\qquad x = 0.4g\ H_2$

5. The electrolysis of molten aluminum oxide results in the production of aluminum and oxygen. (a) Write a balanced equation for the reaction. (b) How many moles of aluminum oxide will be required to produce 10.8 g of aluminum? (c) How many grams of oxygen will be produced if 10.8 g of aluminum are produced?

a) $2Al_2O_3 \rightarrow 4Al + 3O_2$ (g)

b) $\frac{27g}{1\text{ mole Al}} = \frac{10.8g}{x\text{ mole}}$ $\qquad$ $\frac{2\text{ mole } Al_2O_3}{4\text{ mole Al}} = \frac{x\text{ mole } Al_2O_3}{.4\text{ mole Al}}$

$x = .4$ mole Al $\qquad$ $x = .2$ mole Al_2O_3

c) $\frac{1\text{ mole Al}}{2\text{ mole } O_2} = \frac{.4\text{ mole Al}}{x\text{ mole } O_2}$ $\qquad$ $\frac{32g}{1\text{ mole } O_2} = \frac{x\text{ g}}{1.2\text{ mole } O_2}$

$x = 1.2$ mole O_2 $\qquad$ $x = 38.4$g O_2

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

PROBLEM SET e: REACTIONS WITH ONE REACTANT IN EXCESS

1. Ammonia gas (NH_3) can be manufactured by combining hydrogen and nitrogen gases. If 28g of nitrogen and 7g of hydrogen are available, how many grams of ammonia can be produced?

   Step 1. Write a balanced equation for the reaction.

   $3H_2 + N_2 \rightarrow 2NH_3$

   Step 2. Determine how many moles of each substance are available.

   N_2 $\frac{2g}{1\text{ mole}} = \frac{7g}{x\text{ mole}}$ $x = 3.5$ mole N_2 $\frac{28g}{1\text{ mole}} = 1$ mole N_2

   Step 3. Which substance is in excess?

   3.5 mole N_2 - 3 mole N_2 $\rightarrow$ .5 mole N_2 in excess

   Step 4. (a) How many moles of ammonia will be produced? (b) How many grams?

   2 mole of NH_3 = 2(14g + 3g) = 34g NH_3

2. 16.6 g of sodium carbonate reacts with 9.49 g of HCl. Water and carbon dioxide are two of the three products formed. What will be the weight of the third product?

   $Na_2CO_3 + 2HCl \rightarrow H_2O + CO_2 + \underline{2NaCl}$

   16.6g Na_2CO_3 + 9.49g $2HCl \rightarrow H_2O + CO_2 + \underline{2NaCl}$

   2(23g) + 12g + 3(16g) = 106g Na_2CO_3 1g + 35.5g = 36.5g HCl

   $\frac{106g}{1\text{ mole}} = \frac{16.6g}{x\text{ mole}}$ $\frac{36.5g}{1\text{ mole}} = \frac{9.49g}{x\text{ mole}}$ $\frac{1\text{ mole } Na_2CO_3}{2\text{ mole NaCl}} = \frac{0.16\text{ mole } Na_2CO_3}{x}$

   $x = 0.16$ mole $x = 0.26$ mole $x = 0.32$ mole NaCl

   23g + 35.5g = 58.5g/mole NaCl $\frac{58.5g}{1\text{ mole}} = \frac{x\text{ g}}{0.32\text{ mole}}$

   $x = 18.72$g NaCl

INVESTIGATION 10: MASS RELATIONSHIPS

3. 72g of magnesium are reacted with 511.25g of copper(II)chloride. Elemental copper begins forming on the magnesium.

(a) What type of chemical reaction is this? (Write a balanced equation.)

$Mg + CuCl_2 \rightarrow Cu + MgCl_2$ Simple Displacement

(b) Which chemical will be completely used up at the end of the reaction?

$64g + 2(35.5g) = 135g/mole\ CuCl_2$

$$\frac{24g}{mole} = \frac{72g}{x} \qquad \frac{135g}{1\ mole} = \frac{511.25g}{x}$$

$x = 3\ mole\ Mg$ $x = 3.79\ mole\ CuCl_2$

Magnesium. 1 mole Mg reacts with 1 mole $CuCl_2$, therefore there is excess $CuCl_2$.

(c) How many grams of copper will be produced?

$$\frac{1\ mole\ Mg}{1\ mole\ Cu} = \frac{3\ mole\ Mg}{x\ mole\ Cu} \qquad \frac{63.5g}{1\ mole} = \frac{x\ g}{3\ mole}$$

$x = 3\ mole\ Cu$ $x = 190.5g\ Cu$

4. 17g of silver nitrate are mixed with 17g of NaCl. What will be the weight of the precipitate?

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

$108g + 14g + 3(16g) = 170g/mole\ AgNO_3$

$108g + 35.5g = 143.5g/mole\ AgCl$

$$\frac{170g}{1\ mole} = \frac{17g}{x} \qquad \frac{1\ mole\ AgNO_3}{1\ mole\ AgCl} = \frac{0.1\ mole\ AgNO_3}{x}$$

$x = 0.1\ mole\ AgNO_3$ $x = 0.1\ mole\ AgCl$

$$\frac{143g}{1\ mole} = \frac{x}{0.1\ mole\ AgCl}$$

$x = 14.3g$

INVESTIGATION 10: MASS RELATIONSHIPS

5. Given the reaction:

$AlCl_3 + 4NaOH \rightarrow NaAlO_2 + 3NaCl + 2H_2O$

413.85 grams of $AlCl_3$ are reacted with an excess of NaOH. How many grams of $NaAlO_2$ will be produced by this reaction?

$27g + 3(35.5g) = 133.5g/mole\ AlCl_3$ $\quad 23g + 27g + 2(16g) = 82g/mole\ NaAlO_2$

$$\frac{133.5g}{mole} = \frac{413.85g}{x\ mole} \qquad x = 3.1\ mole\ AlCl_3$$

$$\frac{1\ mole\ AlCl_3}{1\ mole\ NaAlO_2} = \frac{3.1\ mole\ AlCl_3}{x} \qquad x = 3.1\ mole\ NaAlO_2$$

$$\frac{82g}{1\ mole} = \frac{x}{3.1\ mole} \qquad x = 254.2g\ NaAlO_2$$

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

TEACHER'S GUIDE

INTRODUCTION

During the Exploration phase of this investigation heat characteristics of systems as related to solution processes are studied. A change in temperature of the water indicates that the substance released or gained heat energy. Examples of exothermic and endothermic processes are investigated. The relationship between the amount of the substance and the amount of heat involved is investigated.

The purpose of the Expansion of the Idea phase is to show that the heat laws apply in chemical as well as physical interactions. Molar heat of solution, molar heat of reaction, and heat capacity as defined in the first reading. Then the students work problems involving molar heat of solution and molar heat of reaction. The second reading introduces the idea of activation energy.

CONCEPT

Reactions involve changes in the total energy of the reacting substances. The amount of energy is proportional to the amount of reacting substances.

EQUIPMENT

A. Exploration

1. Labware

a. 18 x 150 mm test tubes
b. Test tube rack
c. Styrofoam cups (approximately 300 ml size)
d. -10 to 110°C thermometers
e. Plastic spoons

2. Reagents

a. Sodium hydroxide pellets
b. Ammonium nitrate
c. Anhydrous magnesium sulfate
d. Hydrated magnesium sulfate

NOTE: If anhydrous magnesium sulfate is not available, heat hydrated magnesium sulfate in a beaker over an electric hot plate at a high temperature until it is dehydrated. Then, grind the anhydrous magnesium sulfate in a mortar and pestle.)

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

B. EXPANDING THE IDEA: Demonstration and Experiment

DEMONSTRATION

1. Labware

   a. Laboratory burner
   b. Asbestos pad

2. Reagents

   a. Ammonium dichromate--25 g

EXPERIMENT

1. Labware

   a. 18 x 150 mm test tube
   b. Test tube clamp
   c. Laboratory burner

2. Reagents

   a. Ammonium chloride
   b. Calcium hydroxide

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

EXPLORATION

1. Since sodium hydroxide will absorb moisture from the air, the teacher should pass out the sodium hydroxide pellets to the students.

2. through 4. The ammonium nitrate, anhydrous magnesium sulfate, and hydrated magnesium sulfate should be set out in labeled, covered containers. Students should be instructed to put the covers back on the containers after the samples are taken.

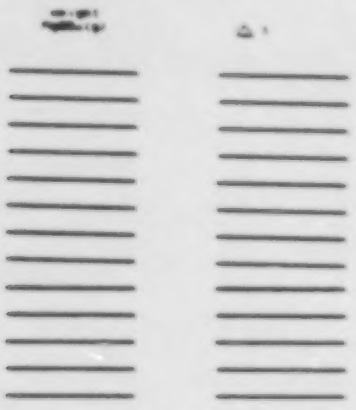
5. To have a range of values, assign various weights of anhydrous magnesium sulfate to the laboratory groups.

6. Caution the students to be careful with the thermometer. A styrofoam cup will not support the thermometer.

7. The students will be using the thermometers as a stirring rod.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING



9. Put this table on the board for the students to record their data. Then they can complete the table of class data on their sheets.

10. The teacher may need to assist the students in setting up the coordinates for the graph. Put the weight on the x-axis and ΔT on the y-axis.

CONCEPTUAL INVENTION (THE IDEA)

Acceptable Answers

1. Some systems get cool when the substance dissolves. Other systems are heated when the substance dissolves.

NOTE: Actually, the answer is when the substance dissolves, it either absorbs heat or gives off heat. This is probably too much to expect of your students.

2. The trend shown by the data on the graph would be that the greater the weight of the magnesium sulfate, the greater the temperature change.

3. This value is obtained by determining the slope of the line ($\Delta T/g$) and multiplying that value by 100 g.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

INVENTION DISCUSSION

A. Ask the students to review the observations they made in the Exploration, Steps 1 through 4. List the substances according to whether the temperature increased or decreased.

B. Discuss what caused the temperature to increase in the sodium hydroxide system.

   1. If the temperature of the system rises when sodium hydroxide is added, what is the source of the heat? (NaOH.)

   2. Was the sodium hydroxide hotter than the water when it was added? (No, it was at the same temperatue.)

   3. What was the source of the energy? (The energy was tied up in the sodium hydroxide.)

C. Ask the students how the energy of the sodium hydroxide in the solution compares with the energy contained in the sodium hydroxide before it was dissolved. (The energy of the sodium hydroxide that is in solution is less.)

D. Now discuss what caused the ammonium nitrate solution and the hydrated magnesium sulfate to cool.

   1. What does that indicate about the total energy of those substances before and after dissolving? (The substances have more energy after dissolving.)

   2. What was the source of the energy gained by these substances? (The water.)

E. Invent the ideas that solution processes that give off heat are called <u>exothermic,</u> and that the heat given off was tied up in the solute prior to the dissolving. Solution processes that absorb heat are called <u>endothermic,</u> and the heat must come from one or more elements of the environment, and is absorbed by the solute.

F. Direct the attention of the students to the graphs of data for the quantitative solution process of anhydrous $MgSO_4$.

1. Ask the students to deduce a relationship between the heat generated and the amount of $MgSO_4$. (This will be a linear, direct relationship.)

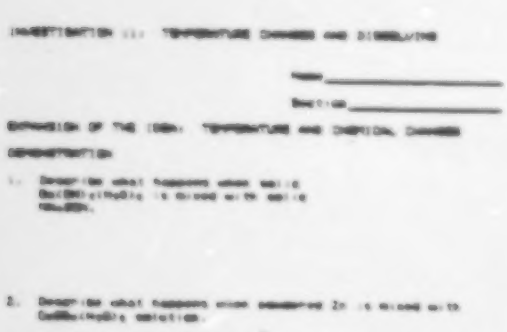
2. Ask if the graph could be used to predict the amount of heat produced by a weight of $MgSO_4$ that would be beyond the range of the present graph. When the class has verbalized its understanding of the idea, summarize the discussion with the following invention statement:

Reactions involve changes in the total energy of the reacting substances. The amount of energy is directly proportional to the amount of reacting substances.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

EXPANSION OF THE IDEA

A. Temperature and Chemical Changes: Demonstration. This activity is important for helping the students understand that the concepts invented during the Exploration phase are true for chemical changes as well as physical changes.



NOTES:

1. Put between 30 and 40 grams of solid NH_4SCN and solid $Ba(OH)_2(H_2O)_8$ in separate beakers. Take the temperature of each solid. Mix the solids together while stirring with the thermometer. Put a little water on a cardboard square, then set the beaker on the cardboard. The cardboard will freeze to the beaker. Ask the students for observations (final temperature and odor).

2. Pour hydrated copper(II)sulfate powder in a beaker. Add water and stir with a thermometer. Record the temperature. Discuss what the temperature of the powdered zinc would be. (Room temperature.) Put 2 teaspoonfuls of powdered zinc in the copper(II)sulfate solution. Stir with a thermometer and record final temperature.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

[illegible]

1. [illegible] evidence to indicate that the above interactions are chemical changes.

2. Write reactions for the reactions.

3. Predict what the effect on the temperature would be if you changed the amounts of chemical reactants in this demonstration.

[illegible]

Questions

1. Temprature change, gas production, two solids reacted to form a liquid.

2. $Ba(OH)_2(H_2O)_8 + 2NH_4SCN \rightarrow Ba(SCN)_2 + 2NH_3 + 10H_2O$

   $CaNH_4(H_2O)_2 + 2I \rightarrow CaI_2 + 2NH_3 + 2H_2O$

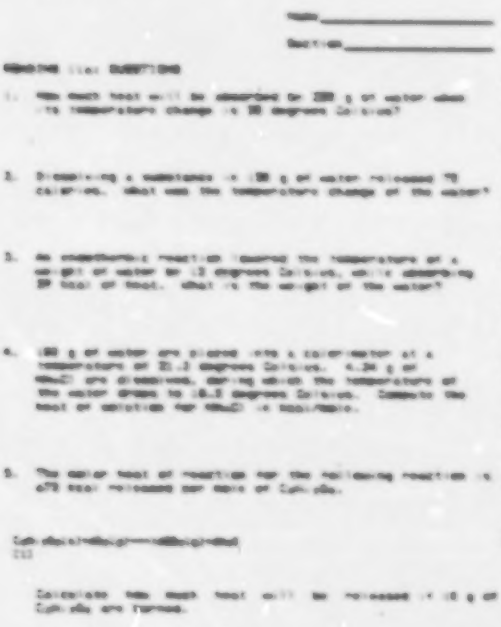
NOTE: The teacher may want to leave the water of hydration out of the reactants. Since water is a product is the first reaction, this will be difficult for the students. A hint that might be given is to ask the students if H_2 was produced. It isn't, and the ammonia gas will not account for all of the hydrogen atoms. Then suggest water as a possible product.

3. If the amount of reactants were increased, the change in temperature would increase proportionally. A decrease in the amount of reactants would decrease the changes in temperature proportionally.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

B. Assign Reading a: MAKING THINGS HOT OR COLD and its questions. This reading concentrates on molar heat of solution, molar heat of reaction, and heat capacity. Even though the reading discusses heat capacity in general, the students are not asked to work with it.



Acceptable Answers

1. $Q = m \triangle t°C = (250\ g)(25°C)(1\ cal/g°) = 6750\ cal$

2. $Q = m \triangle t°C$

$\triangle t = Q/mc = 75\ cal/(150\ g)(1\ cal/g°C)$

3. $Q = m \triangle t$

$m = Q/\triangle t°C = 39000\ cal/(13°C)(1\ cal/g°C) = 3000\ g$

4. Heat = $m(\triangle t)C$

$100\ g\ (21.3°C - 18.5°C)(1\ cal/g°C)$

$(100\ g)(2.8°C)(1cal/g°C) = 280\ cal$

5. $(280\ cal/4.34\ g)(53.5\ g/mole)(1\ kcal/1000\ cal) = 34.5\ kcal/mole$

$x/10\ g = (673\ kcal/1\ mole)(1\ mole/100\ g)$

$x = 37.39\ kcal$

During the post-reading discussion, the teacher might want to give the students the equation and units for heat capacity. The equation for heat capacity is $Q = m(\triangle t)C$, where m equals the weight in grams, C equals the heat capacity (cal/g°C), and $\triangle t$ equals the change in temperature.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

C. INPUT HEAT: DEMONSTRATION

[illegible]

1. [illegible] and heats $(NH_4)_2Cr_2O_7$.

2. [illegible]

3. [illegible]

4. [illegible]

[illegible]

1. [illegible]

2. [illegible]

1. Pour two teaspoonfuls $(NH_4)_2Cr_2O_7$ on an asbestos square. Light with a burner. (Be ready to clean up a mess.)

2. NH_4Cl and $Ca(OH)_2$ should be in covered containers.

3. Spread the solids out along the side of the test tube. Heat gently, moving in and out of the flame. Stress that they make careful observations.

1. The reaction in Step 1 was a chemical change. The evidences for a chemical change were heat, light, color change and gas production.

In Step 2 the observations were moisture collecting at the top of the test tube and ammonia gas being produced. It was a chemical change.

2. The first reaction was exothermic and the second reaction was endothermic.

Discussion: Compare the energy of the products with the energy of the reactants. To assist the students in understanding that the $Ca(OH)_2 + NH_4Cl$ reaction is endothermic use the observation that the production of ammonia gas stops when the system is not being heated.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

D. Assign Reading b: GREMLINS AND OILY RAGS and its Questions. A discussion of the significant points in the reading and the questions should be held.

1. Point out to the students that the spun fiberglass is like the fiberglass bats that are used for attic insulation.

The spun fiberglass has dead air spaces. Since dead air spaces are poor conductors of heat, the fiberglass is an insulator. Plate glass does not have dead air spaces, so it is a better conductor than fiberglass. However, plate glass is not a good conductor of heat either.

2. The students will probably need to be reminded of the investigations in which they studied these reactions—Investigation 9, Decomposition of Hydrogen Peroxide; and Investigation 8, Zinc and Hydrochloric Acid.

At room temperature, zinc and hydrochloric acid have a low activation energy. The decomposition of hydrogen peroxide has a higher activation energy.

3. Heating the system will add energy to the particles. The more energy the reacting particles have, the greater the chance that they will react when they collide.

BEST COPY AVAILABLE

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

DISCUSSION

The reading defines activation energy as the minimum amount of energy that is required in a system before a chemical reaction will occur. The students will need to discuss this idea. Ask the students to think about systems they have studied.

1. We start with the total energy of the reactants.

2. Either from the environment or from an outside source of energy, the energy is increased to a point that is called the activation energy.

3. In some cases, like zinc and hydrochloric acid, the reactants have enough energy at room temperture to be at or above activiation energy.

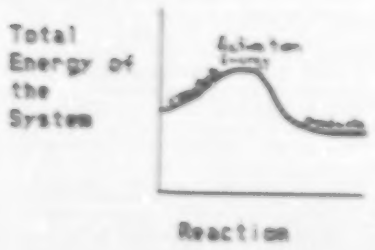
4. In other systems, such as the calcium hydroxide and ammonium chloride system, energy has to be added to the reactants to get them to the minimum energy need to react.

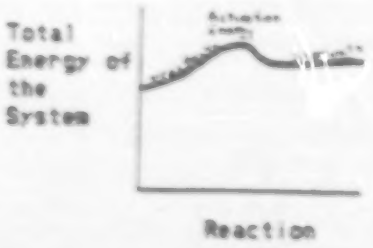
5. At this peak point in the energy of the system, some energy is given off.

   a. If the reaction is endothermic, some energy is absorbed by the system. The energy level will go down, but not as low as it was originally.

   b. If the reaction is exothermic, the energy will go down farther than it was originally.

6. This information can be shown graphically.





The horizontal coordinate can be confusing to the students. They will want to know what it is. It is a generalized coordinate. The graph is to show the difference between the total energy of the reactants and the total energy of the products.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

TEACHER'S GUIDE

INTRODUCTION

In this Investigation the student will be applying a model for ions formulated during Investigation 6. The model will be extended through laboratory exploration to eighteen binary substances. Primarily, the experience will lead to conceptual inventions concerning the role of ions in the conductivity of solutions and the dissolving of ionic substances in water. The idea of acidity and basicity of some solutions is introduced in this investigtion, but it will not be completely resolved until Investigation 13: Arrhenius Solutions.

CONCEPT

Soluble ionic substances dissolve because with the ions of the solid causing the ions to separate and mix homogeneously with the water.

EQUIPMENT

A. Exploration

1. Labware

a. Plastic spoons
b. 150 ml beakers
c. Glass stirring rods
d. Ring stands
e. Conductivity apparatus
f. Pink and blue litmus paper

2. Reagents

Approximately 100 g each (in covered dishes) of:

a. Iron (II) sulfide
b. Barium chloride
c. Nickel (II) chloride
d. Sodium fluoride
e. Calcium chloride
f. Copper (II) oxide
g. Aluminum chloride
h. Cobalt (II) chloride
i. Potassium iodide
j. Chromium (III) chloride
k. Sodium bromide
l. Magnesium chloride
m. Lead oxide

n. Calcium oxide
o. Zinc sulfide
p. Iron (III) oxide
q. Strontium chloride
r. Potassium bromide

NOTE: The solids to be used in this Investigation can be conveniently placed in petric dishes with the written names of the compounds on the dish covers. This will provide further practice in recognizing both the formula and the written name of a substance.

s. Distilled water in wash bottles

B. Expanding the Idea (Demonstration)

1. Labware

a. 500 ml florence flask and stopper
b. Laboratory burner

2. Reagents

a. Sodium acetate solid--150 g
b. Distilled water

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

EXPLORATION

[redacted]

1. Two periods will probably be required to perform the tests and answer THE IBDA questions.

2. It is important for the students to perform these tests thoughtfully.

Students should be reminded to rinse the conductivity apparatus electrodes between uses.

Tell the students that conductivity tests should be performed whether all of the solid dissolves or not.

[redacted]

BEST COPY AVAILABLE

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

CONCEPTUAL INVENTION (THE IDEA)

[illegible]

Acceptable Answers

1. (a) If the solid disappeared, it dissolved.
(b) If the solid partially disappeared, and it conducted and/or changed litmus paper, the solid partially dissolved.

2. In general, the crystalline substances are soluble and the powders are insoluble. Exceptions to this statement include FeS, MgF, and CaS. Color and solubility appear to be correlated.

There are not any other patterns between physical appearance of the substance and their solubilities.

NOTE: If the students list artificial patterns, there is no real harm done.

3. When a substance dissolves it breaks down into smaller pieces to form a homogeneous mixture.

BEST COPY AVAILABLE

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

INVENTION DISCUSSION

A. The initial discussion should center around the means by which solubility was determined. When the entire class has accepted a set of criteria for determining solubility, ask the class to determine what characteristic (or characteristics) all the soluble compounds had in common. This common characteristic is conductivity. The following line of questioning may be helpful in leading students to the above conclusion.

1. How did you determine whether a compound was insoluble or soluble?

2. How could you tell if the solid partially dissolved?

3. About which substances did you have questions as to whether they totally dissolved?

4. Let's discuss one of the substances listed in Question 3.

   a. How could you tell if this substance partially dissolved?
   b. Was a solution formed? Explain.
   c. What did the reactions to litmus paper indicate?

5. What is a solution?

6. What are the identifiable characteristics of a solution?

7. Which substances did you find to be insoluble? What evidence do you have to support this decision?

8. What characteristic(s) do all the soluble compounds have in common?

B. Review, by way of questioning, what happens when a substance conducts electricity. (A good reference to use is the electroscope demonstration in Investigation 6.)

1. When a substance conducts electricity, it allows an electric current to flow.

2. Electric current is the movement of electrical charge.

C. Why doesn't distilled water conduct an electrical current? (There are no charges that are free to move in the water.)

NOTE: The relatively small amount of ionization of water is not important here. In fact, introducing the ionization of water at this time would probably be confusing. The ionization of water will be discussed during Investigation 13.

D. In light of the reasoning about water's lack of conductivity, ask the class to suggest why a solution of KCl, for example, <u>does</u> conduct electricity. Guide the student input through interaction and questioning to get to the idea that the solution must contain charged particles that are free to move.

E. Now we are reading to answer the question that was originally asked of the students concerning what happens to the compounds when they are dissolved.

1. According to the ionic model for compounds, such as KCl, ask the class what a crystal of KCl would contain (positive potssium ions and negative chloride ions).

2. An example diagram could be drawn on the board to help get to the idea that each ion in the crystal is attracted by several others. That is, of course, an oversimplified picture, but sufficient for this discussion. Stress the three dimensionality of the structure.

3. Ask how water would need to break this crystal down during dissolving in order to allow the solution to contain charged particles that are free to move. Be prepared to accept a variety of responses. Work with the students to reach the point that the crystal is pulled apart into ions. This leads to the conceptual invention: Soluble ionic substances dissolve because water interacts with the ions of the solid causing the ions to separate and mix homogeneously with the water.

4. The teacher will have to decide how detailed a model for solution conductivity is desirable, but the most important understanding for the students to grasp is that a solution containing ions can move will allow for electrical conductivity.

F. A good question to follow up the ideas discussed here would be one such as: "Why does KCl dissolve in water but PbO does not?" A good response might be: "Water can overcome the attraction between potassium and chlorine, but not between lead and oxygen." The point should be clearly made that water is able to exert enough force on the soluble solids to pull them apart and hold them in solution.

G. At this point, you may wish to have the students refine their definition of an electrolyte by adding the cause of solution conductivity. (The first definition was made in Investigation 1.)

EXPANSION OF THE IDEA

A. Engage in a discussion comparing the CaO system to one of the others that totally dissolved.

1. It will be essential to have established that some CaO dissolved. (This probably has already been done in INVENTION DISCUSSION.)

2. Establish, through questioning, that the essential difference between the two systems is the extent to which the dissolving process takes place.

3. The model for dissolving can be brought into the discussion with relation to the attraction between the ions in the compounds. One model might be that more water molecules are required to separate the ions of Ca^{2+} and O^{2-} than are required in the other substance. If this is valid, the water available for dissolving would be used up earlier in the CaO system than in the other system.

4. Invent <u>saturated solution</u> as the term applied to a solution that has dissolved all the solute it can hold in homogeneous mixture.

5. Ask the students to suggest factors that would affect the amount of solute that would dissolve (for example, the volume of solvent used and the temperature of the solvent).

B. Assign Reading 12, H_2O vs. THE SOLUTE and QUESTIONS. This reading can be assigned as homework. The importance of the material will warrant at least one period of discussion. Acceptable responses to the questions are listed below.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Name_____

Section_____

READING 12: QUESTIONS

1. Referring to Table I of the Reading, state the general trend in solubility of solids as the temperature of the solvent is increased.

2. Referring to Table I of the Reading, state the general trend in solubility of gases as the temperature of the solvent is increased.

3. Suggest reasons for each of the above-stated trends. (Hint: Temperature is a measure of the average kinetic energy of the molecules. Higher temperatures, then, indicate higher kinetic energy.)

175

Acceptable Answers

1. The solubility of solids tends to increase with increasing solvent temperature. One exception to this tendency, listed in Solubility Table 1, is calcium hydroxide.

2. Gases tend to be less soluble as solvent temperature increases.

3. Comparing original states of the solutes to the state toward which they tend with increasing solvent temperature allows for some explanation of the patterns. The higher the kinetic energy of the system, the faster the particles move. They tend toward a more random state. As the motion increases, the attractive forces become less effective at holding the solvent and solute together (gas-liquid system) and the gas comes out of solution. In the solid-liquid system the solid's attractive forces are partially overcome by increased particle motion, so the attractive forces from the solvent particles can be more effective in separating the solid particles.

BEST COPY AVAILABLE

4. Make a study of Solubility Table 2 and use that information to fill in the following chart.

Ion	Soluble or Insoluble (Pattern)	Exceptions to the Pattern
Ammonium	soluble	none
Acetate	soluble	tin (II)
Bromide	soluble	mercury (I), silver
Chloride	soluble	mercury (I), silver
Iodide	soluble	mercury (I) and (II), silver
Oxide	insoluble	barium, calcium, potassium, sodium, strontium
Sulfate	soluble	barium, lead, mercury (II)
Nitrate	soluble	tin (II)
Carbonate	insoluble	ammonium, potassium, sodium
Phosphate	insoluble	ammonium, hydrogen, magnesium, potassium, sodium
Potassium	soluble	none
Sodium	soluble	none

BEST COPY AVAILABLE

5. Given the following sets of reactants, predict the products, balance the equations, and use the patterns from Question 4 to predict whether a precipitate will form.

(a) $3NH_4OH + Al(C_2H_3O_2)_3 \rightarrow Al(OH)_3 + 3NH_4C_2H_3O_2$

Precipitate: $Al(OH)_3$

(b) $2LiClO_4 + MgSO_4 \rightarrow Li_2SO_4 + Mg(ClO_4)_2$

No precipitate.

(c) $AgNO_3 + KI \rightarrow KNO_3 + AgI$

Precipitate: AgI

(d) $3Pb(NO_3)_2 + 2H_3PO_4 \rightarrow 6HNO_3 + Pb_3(PO_4)_2$

Precipitate: $Pb_3(PO_4)_2$

(e) $Ba(OH)_2 + CuSO_4 \rightarrow BaSO_4 + Cu(OH)_2$

Precipitate: Both products.

Discussion of the Reading 12 Questions can be followed by questions relating to solution concentration, such as:

1. "If two solutions of different substances in water each contained 100 g of solute, how would the concentration of those solutions compare?"

2. "What is meant by solution concentration?"

3. "How could we know that two different solutions had the same concentration?"

The aforestated questions can be used in discussion to lead to a definition (teacher invention) of <u>molarity</u> (number of moles of solute per liter of solution).

Work through some examples with the students to show how this concentration unit can be determined and used. Examples follow:

1. 3.31 g $Pb(NO_3)_2$ in enough water to make 500 ml of solution. (0.1 mole/0.5 liter = 0.02 mole/liter.)

2. What is the molar concentration if 55.5 g $CaCl_2$ were added to 1 liter of water? (Unknown, because solution volume would be unknown.)

3. 75 g of $CaCO_3$ are reacted with HCl to produce CO_2 and two other products. How many ml of 3M HCl would be required to react all the $CaCO_3$?
$CaCO_3 + 2HCl \rightarrow CO_2 + CaCl_2 + H_2O$

$$(75\text{ g})\ \frac{1\text{ mole}}{100\text{ g}}$$

$$\frac{1\text{ mole } CaCO_3}{2\text{ mole HCl}} = \frac{0.75\text{ mole } CaCO_3}{x\text{ mole HCl}}$$

$x = 0.75$ mole $CaCO_3$ $\qquad x = 1.5$ mole HCl

$$\frac{1\text{ liter HCl}}{3\text{ mole HCl}} = \frac{x\text{ liter HCl}}{1.5\text{ mole HCl}}$$

$x = 0.5$ liter HCl solution

C. Assign PROBLEM SET 12: SOLUTION CONCENTRATION EXERCISES.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

EXPANSION OF THE IDEA

PROBLEM SET 1a: SOLUTION CONCENTRATION EXERCISES

1. Calculate the molarity of 1.5 liters of a solution of $KClO_3$ containing 24.5 g of the solute.

$$(24.5\ g)\ \frac{1\ mole}{122.5\ g} = 0.2\ moles\ KClO_3$$

$$\frac{0.2\ moles}{1.5\ liter} = 0.13\ moles/liter = 0.13M$$

2. What amount of ammonium dichromate would be needed to add to sufficient water to make one liter of 0.5M solution?

$(NH_4)_2Cr_2O_7 = 252 g/mole$

$$\left(\frac{252\ g}{1\ mole}\right)\left(\frac{0.5\ mole}{1\ liter}\right) = 126\ g/liter$$

3. Calculate the molarity of 0.17 g of barium hydroxide in 250 ml of solution.

$Ba(OH)_2 = 171\ g/mole$

$$(0.17\ g)\ \frac{1\ mole}{171\ g} = 0.001\ mole$$

$$\frac{0.001\ mole}{0.25\ liter} = 0.004\ \frac{mole}{liter} = 0.004M$$

4. How would you prepare 500 ml of a 0.7M solution of sodium hydroxide?

NaOH = 40 g/mole
0.35 moles NaOH are needed for 500 ml of 0.7M solution.

$$(0.35\ mole)\ \frac{40\ g}{mole} = 14\ g$$

Dissolve 14 g of NaOH in enough water to make 500 ml of solution.

5. How much KCl would you need to prepare 2.5 liters of 0.01 molar solution?

KCl = 74.5 g/mole 0.01M = 0.745 g/liter

$\frac{0.745\ g}{1\ liter}$ (2.5 liter) = 1.86 g

6. What is the molarity of acetic acid in vinegar which contains 5.0 grams of acetic acid per 100 ml of vinegar?

$HC_2H_3O_2$ = 60 g/mole

$\frac{5.0\ g}{100\ ml} = \frac{x\ g}{1000\ ml}$; x g = 50 g/liter

$\frac{50\ g}{liter} \times \frac{1\ mole}{60\ g}$ = 0.83 mole/liter = 0.83M

7. If 200 ml of a certain solution contains 28.2 grams of magnesium chloride, what is its molarity?

$MgCl_2$ = 95.3 g/mole

(28.2 g) $\frac{1\ mole}{95.3\ g}$ = 0.3 mole

$\frac{0.3\ mole}{200\ ml} = \frac{x\ mole}{1000\ ml}$; x mole = 1.5 mole or 1.5M

8. How many grams of solute would be contained in 5 liters of 0.25M sodium chloride solution?

NaCl = 58.5 g/mole

$\frac{0.25\ mole}{1\ liter} = \frac{x\ mole}{5\ liter}$; x = 1.25 mole

(1.25 mole)(58.5 g/mole) = 73.1g

9. How many grams of solute would be contained in 450 ml of a 0.5M solution of lithium perchlorate?

$LiClO_4$ = 106.5 g/mole

solution contains 53.25 g/liter

$$\frac{53.25\ g}{1000\ ml} = \frac{x\ g}{450\ ml} \; ; \; x\ g = 24\ g$$

10. Formaldehyde (CH_2) is a gas at room temperature. Its maximum solubility in water at room temperature and atmospheric pressure is 39.58 g per 100 ml of solution. What is the molar concentration of such a solution?

CH_2O = 30 g/mole

$$\frac{(395.8\ g)}{(1\ liter)} \; \frac{(1\ mole)}{(30\ g)} = 13.2\ mole/liter = 13.2M$$

EXPANSION OF THE IDEA

PROBLEM SET b: SOLUTION CONCENTRATION EXERCISES

1. What is the molar concentration of 21.3 g of lithium perchlorate dissolved in enough water to make one liter of solution?

$$(21.3\ g)\ \frac{1\ mole}{106.5 g} = 0.2\ mole \qquad M = \frac{0.2\ mole}{1\ liter} = 0.2M$$

2. What volume of water would have to be added to the solution in Exercise 1 to make the concentration 0.001M?

$$(0.2\ mole)\ \frac{0.001\ mole}{1\ liter} = 200\ liter \qquad -\ \frac{200\ liter}{1\ liter}$$

$$199\ liter$$

3. Given that the solubility of $Ca(OH)_2$ in water is 0.18 g/100 ml, what would be the molar concentration of a saturated water solution of $Ca(OH)_2$?

$$\frac{0.18\ g}{100\ ml}(10) = \frac{1.8\ g}{1000\ ml}\ ;\ (1.8\ g)\ \frac{1\ mole}{74\ g} = 0.24\ mole$$

$$M = \frac{0.024\ mole}{1\ liter} = 0.024M$$

4. How many grams of solute would be contained in 250 ml of a 6M NaOH solution?

$$(250\ ml)\ \frac{6\ moles}{1000\ ml} = 1.5\ moles;$$

$$(1.5\ moles)\ \frac{40\ g}{1\ mole} = 60\ g\ NaOH$$

5. You know that zinc reacts with hydrochloric acid to produce hydrogen gas and zinc chloride. How many milliliters of a 4M HCl solution would be required to produce 1 mole of hydrogen gas, using this reaction?

$Zn + 2HCl \rightarrow H_2 + ZnCl_2$

$$\frac{2 \text{ mole HCl}}{1 \text{ mole } H_2} = \frac{x}{1 \text{ mole } H_2}; \quad x = 2 \text{ mole HCl}$$

$$\frac{4 \text{ mole HCl}}{1000 \text{ ml}} = \frac{2 \text{ mole}}{x \text{ ml}}; \quad x = 500 \text{ ml HCl solution}$$

6. The solubility of lead iodide is 0.05 g/100 ml of water. If 0.02 mole of lead nitrate is reacted with 0.04 mole of potassium iodide, what will be the minimum volume of solution required to keep a precipitate from forming?

$Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$

0.05 g/100 ml = 0.5 g/1000 ml

$$(0.5 \text{ g}) \frac{1 \text{ mole}}{461 \text{ g}} = 0.001 \text{ mole } PbI_2$$

$$\frac{1 \text{ mole } Pb(NO_3)_2}{1 \text{ mole } PbI_2} = \frac{0.02 \text{ mole } Pb(NO_3)_2}{x}$$

$$x = 0.02 \text{ mole } PbI_2$$

$$\frac{0.001 \text{ mole}}{1 \text{ liter}} = \frac{0.02 \text{ mole}}{x}; \quad x = 20 \text{ liter}$$

7. A water soluble, liquid compound has a density of 0.8 g/ml and a formula of CH_4O. What volume of this liquid would have to be added to water to make 1 liter of a 1.5M solution?

$CH_4O = 12 \text{ g} + 4(1 \text{ g}) + 16 \text{ g} = 32 \text{ g/mole}$

$$\left(\frac{32 \text{ g}}{\text{mole}}\right)\left(\frac{\text{ml}}{0.8 \text{ g}}\right) = 40 \text{ ml/mole}; \quad (1.5 \text{ mole}) \frac{40 \text{ ml}}{\text{mole}} = 60 \text{ ml}$$

8. The solubility of potassium iodide is 128 g/100 ml of water. Is it possible to make a water solution of KI with a concentration of 8M? Explain the reasoning that lead to your response.

KI = 166 g/mole

$$8M = \frac{8 \text{ mole}}{1 \text{ liter}} = \frac{8(166 \text{ g})}{1 \text{ liter}} = \frac{1328 \text{ g}}{1 \text{ liter}}$$

$$\left(\frac{128 \text{ g}}{100 \text{ ml}}\right)\left(\frac{1000 \text{ ml}}{1 \text{ liter}}\right) = \frac{1280 \text{ g}}{1 \text{ liter}}$$

Only 1280 g KI will dissolve in 1 liter of solution; to make an 8M solution you would need 1328 g to dissolve in 1 liter of solution.

9. Explain how you would make a 0.01M solution of $CrCl_2$, using 61.5 g of the solid and all the water you needed.

$CrCl_2$ = 123 g/mole

$$(61.5 \text{ g})\ \frac{1 \text{ mole}}{123 \text{ g}} = 0.05 \text{ mole } CrCl_2$$

$$(0.05 \text{ mole})\ \frac{1 \text{ liter}}{0.01 \text{ mole}} = 50 \text{ liters}$$

10. Suppose that you are given a solution of silver nitrate containing 17 g $AgNO_3$ per liter of solution, and you are given a sodium chloride solution containing 11.7 g NaCl per liter of solution. How many liters of each solution would be needed to produce 0.6 moles of AgCl?

$$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$$

$$(17\ g)\ \frac{1\ mole}{170\ g} = 0.1\ mole\ AgNO_3$$

$$\frac{0.1\ mole}{1\ liter} = \frac{0.6\ mole}{x}\ ;\quad x = 6\ liters\ AgNO_3$$

$$(11.7\ g)\ \frac{1\ mole}{58.5\ g} = 0.2\ mole\ NaCl$$

$$\frac{0.2\ mole}{1\ liter} = \frac{0.6\ moles}{x}\ ;\quad x = 3\ liter\ NaCl$$

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

DEMONSTRATION

The idea of supersaturation of solutions can be introduced with a brief demonstration using supersaturated $NaC_2H_3O_2$ and/or $Na_2S_2O_3$ solutions. The guidelines for the $NaC_2H_3O_2$ demonstration are as follows.

SOLUTION PREPARATION

Mix 150 g of $NaCH_3COO * 3H_2O(s)$ with 40 ml of distilled water in a clean 500 ml Florence flask. Heat the mixture in a burner flame with occasional swirling to dissolve all the sodium acetate. Allow the supersaturated solution to cool to room temperature. The flask should be stoppered to prevent dust from seeding the solution during storage. If the flask is clean and free of scratches, the solution should remain stable for months.

TO CLEAN UP

The solid $NaCH_3COO * 3H_2O$ should be put in the flask. Heat to reform the supersaturated solution. Wash the sides of the flask using a small amount of distilled water. It is important that no crystals remain at the mouth of the flask, since immediate crystallization will occur when you attempt to pour the solution from the flask the next time the demonstration is attempted.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

TEACHER'S GUIDE

INTRODUCTION

This Investigation is designed to develop the idea of Arrhenius acids and bases. The written formulas of acids and bases will be investigated in the Exploration phase. During the Expansion of the Idea phase, the students will explore acids of different strengths, the difference between the concentration and strength of solutions, and dynamic equilibrium. The ideas of pH, acid-base neutralization, Bronsted-Lowry acid and base definitions, chemical equilibrium, and definition of acid-forming and base-forming salts will be introduced in later investigations.

CONCEPT

Arrhenius acids are compounds that release H^+ ions in water solution; bases are compounds that release OH^- ions in water solution.

EQUIPMENT

A. Exploration

1. Labware

a. Plastic spoons
b. 50ml beakers
c. 400ml beakers
d. Light bulb conductivity apparatus
e. Litmus paper (blue and pink)

2. Reagents

a. Solutions and solids for the Exploration should be centrally located in the laboratory. Each should be placed in 400ml beakers and these beakers should be kept covered except when tests are being performed on them. 50g to 100g of solid per beaker will be sufficient. The solutions need only be sufficiently concentrated to give positive litmus and conductivity tests.

Steps 1 and 2

Solids--oxalic acid, magnesium oxide (labeled magnesium hydroxide)
Solutions--hydrochloric acid, sodium hydroxide

Step 4

Solids--boric acid, barium hydroxide, calcium hydroxide
Solutions--acetic acid, nitric acid, ammonium hydroxide, potassium hydroxide, sulfuric acid

B. Expansion of the Idea

1. Labware

a. Eudiometer
b. 25 x 150mm test tube
c. #4 one-hole rubber stopper
d. Right angle glass bend
e. U-bend glass tube or small rubber tubing to extend from the right angle glass tubing to the eudiometer
f. 12 inch length rubber tubing
g. 600ml beaker
h. Ring stand with clamp for eudiometer
i. 10ml graduated cylinder

2. Reagents

a. Solutions of HCl--0.1M, 0.2M, 0.3M, 0.4M
b. Solutions of $HC_2H_3O_2$--0.1M, 0.2M, 0.3M, 0.4M

INVESTIGATION 13: ARRHENIUS SOLUTIONS

EXPLORATION

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Name_____

Section_____

INTRODUCTION

You have noticed that certain of the compounds that you have tested in water solution have changed the color of litmus paper. What designation have we given to solutions that changed litmus paper from blue to pink?_____ What designation have we given to solutions that changed litmus paper from pink to blue?_____

In this investigation you will have the opportunity to begin studying the source of the acid and base characteristics of solutions.

EXPLORATION

1. In the laboratory there are two solutions and two solids in labeled containers. Be sure that you are careful to avoid contaminating the solutions and solids when you are performing tests on them. Test each compound as a solid and as a solution with litmus paper and for conductivity.

2. Classify the compounds into acids and bases.

Acids Bases

3. Look carefully at the formulas for the acids and the bases. Make a hypothesis of the characteristics that make a substance an acid or a base.

4. Using your hypothesis, predict whether the following compounds are acidic or basic. List only your predictions before proceeding to Step 5.

Table I

Compound	Prediction	Results of Conductivity Tests	Reaction to Litmus Paper
HNO_3			
$NaNO_3$			
CO_2			
$Ca(OH)_2$			
CH_4			
H_2SO_4			
$NaCl$			
$Ca(NO_3)_2$			
HCl			

5. In Table I you made predictions about which compounds were acids and which were bases. You may now test those predictions in the laboratory. List the tests you perform below and the results in Table I.

1. The teacher should instruct the students in the proper use of the solutions and solids for Steps 2 and 4.

   a. The beakers of solids and solutions may be taken to the students' laboratory station.

   b. When the students make solutions of the solids, they should use about half a spoonful of solid in their own beakers.

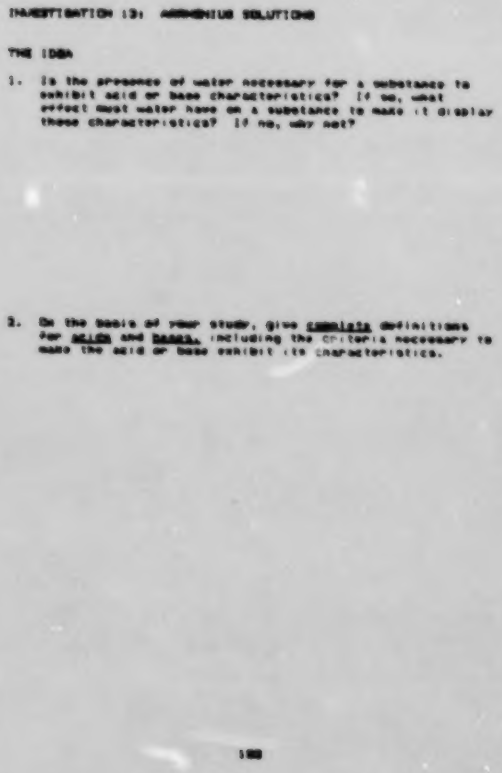
3. After each conductivity test, the electrodes should be rinsed into the sink, not into the solution or solid.

For the Exploration phase, allow two class periods to make the predictions and perform the tests.

BEST COPY AVAILABLE

INVESTIGATION 13: ARRHENIUS SOLUTIONS

CONCEPTUAL INVENTION (THE IDEA)



INVESTIGATION 13: ARRHENIUS SOLUTIONS

THE IDEA

1. Is the presence of water necessary for a substance to exhibit acid or base characteristics? If no, what effect must water have on a substance to make it display these characteristics? If no, why not?

2. On the basis of your study, give complete definitions for ACIDS and BASES, including the criteria necessary to make the acid or base exhibit its characteristics.

188

Acceptable Answers

The responses to the questions should contain the following ideas:

1. The presence of water is essential to the exhibition of acid and base characteristics. The effect of the water is to break the acids and bases into ions, and these ions interact with the indicator in the litmus paper. (It should be noted that many students will not have made the connection between the indication and the type of ion responsible.)

2. (a) An acid is a substance containing hydrogen, which dissolves and ionizes in water to produce hydrogen ions. The water solution will conduct electricity and will change litmus paper from blue to pink.

(b) A base is a substance containing hydroxide, which dissolves and ionizes in water to produce hydroxide ions. The water solution will conduct electricity and will change litmus paper from pink to blue.

INVENTION DISCUSSION

A. In this discussion the teacher will begin with the student responses to the two questions and work toward the development of the Arrhenius definitions for acids and bases.

   1. Ask for student responses to Question 1. The attention of the students should be focused on the ionization process (as evidence by conductivity tests). Students may need to be reminded that water seems to exhibit both positive and negative electrical attraction characteristics (see Investigation 12).

BEST COPY AVAILABLE

BEST COPY AVAILABLE

2. The teacher can ask for volunteers to give their responses to Question 2. The frst people to answer may omit some important factors, but these can be filled in by other student additions. List the definitions on the board.

   Lead the discussion to the idea that the acid solutions have hydrogen ions in common and that basic solutions have hydoxide ions in common.

3. The criteria proposed within Question 2 may be drawn from the written exercises and laboratory findings in the Exploration phase. At least the following points are expected to surface during class interaction:

   a. The substance must be in an aqueous solution.
   b. The substance must ionize in solution (usually conducting electricity).
   c. The solution must change litmus paper.

B. A question should arise as to the nonconductive behavior of boric acid. The teacher may wish to heat a saturated solution to test with a conductivity apparatus. Since the light will glow dimly, its extent of ionization may be discussed in contrast with the much greater ionization of hydrochloric acid. The acidic indication of boric acid to litmus paper should also be pointed out.

C. The conceptual invention is that Arrhenius acids are compounds containing hydrogen which release H^+ ions in a water solution. Bases are compounds containing OH which release OH^- ions in a water solution.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

EXPANSION OF THE IDEA

A. EXPERIMENT

EXPANSION OF THE IDEA: METALS IN ACIDS

EXPERIMENT

1. Write the chemical equation for the reaction of magnesium and hydrochloric acid.

2. What type of chemical reaction is this?

3. Set up your laboratory apparatus to collect a gas by displacement of water. You will read the volume of gas collected after each minute for a period of five minutes. First, add 10ml of the acid to the test tube. When you are ready to collect the gas, add a 40cm length of magnesium ribbon to the test tube. Each group will test three different acid solutions according to your teacher's instructions.

4. In Table 2 record the solutions your group will be testing and the data you collect by following the procedure from Step 3.

TABLE I

Kind of Solution	VOLUME OF GAS (in ml) AFTER				
	1 min.	2 min.	3 min.	4 min.	5 min.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

5. Record the class data in Table 3.

TABLE I

VOLUME OF GAS (in ml) AFTER

ACID SOLUTION	1 min.	2 min.	3 min.	4 min.	5 min.
1.0 HCl					
1.0 H_2SO_4					
1.0 HCl					
1.0 H_2SO_4					
1.0 HCl					
1.0 H_2SO_4					
1.0 HCl					
1.0 H_2SO_4					

1. $Mg + 2HCl \rightarrow MgCl_2 + H_2$

2. Single displacement.

3. Each pair of students should be assigned three solutions to test. They should record their solution in Table 13-2, then the class data will be gathered in Table 13-3. The teacher can have the class graph the average values for each solution.

The purpose of measuring the volume of gas produced is to have a quantitative way of observing the relative vigor of the different reactions.

BEST COPY AVAILABLE

TABLE 13-3

VOLUME OF GAS (in ml) AFTER

ACID SOLUTION	1 min.	2 min.	3 min.	4 min.	5 min.
0.1M HCl	8.7	13	14.8	15.7	16.2
average					
0.1M $HC_2H_3O_2$	2	3	3.5	4	4.3
average					
0.2M HCl	14	21.2	24	25.2	25.9
average					
0.2M $HC_2H_3O_2$	9.7	15.7	18.7	20.4	21.4
average					
0.3M HCl	27	35	37	38.1	39
average					
0.3M $HC_2H_3O_2$	14.5	25	30	32.2	34.2
average					
0.4M HCl	37.5	46	48.2	49	49.6
average					
0.4M $HC_2H_3O_2$	15	28	34	37.5	39.7
average					

BEST COPY AVAILABLE

INVESTIGATION 13: ARRHENIUS SOLUTIONS

INVESTIGATION 13: ARRHENIUS SOLUTIONS

QUESTIONS

1. Does the concentration of the acid solution affect the reaction rate? Explain.

2. Compare the reaction of magnesium with 0.4M HCl to the reaction of magnesium with 0.4M $HC_2H_3O_2$. Offer an explanation for any differences in the rates of the reactions.

Acceptable Answers

1. The more concentrated the acid solution is the faster the reaction occurs.

2. The reaction rate of Mg with 0.4M HCl was greater than the reaction rate of Mg with 0.4M $HC_2H_3O_2$.

BEST COPY AVAILABLE

Discuss the collected data showing different rates of the same concentration of different acids.

1. Ask the students the difference between a 0.1M HCl solution and a .4M HCl solution. Have the students state what each concentration means. Relate these statements to the effect of different concentrations of H^+ on the relative rates of reaction.

2. The laboratory exercise in Step 4 introduces the students to acids of different strengths. The question for discussion is, "If 1M HCl and 1M $HC_2H_3O_2$ are the same concentration, how can we account for more activity in 1M HCl"? (The difference relates to the extent to which the acids are ionized in the solution.)

B. Assign READING 13a: IONS IN EQUILIBRIUM. The reading should help clarify the reasons for boric acid reacting to litmus paper but not conducting. The students are presented with more information on dynamic equilibrium.

C. Assign READING 13b: THOSE MIGHTY ACIDS, and the questions.

In READING 13b: THOSE MIGHTY ACIDS, the students are given a table of K_A values for several acids. This exercise is to point out, using laboratory evidence, the difference between the concentration and the strength of solutions.

BEST COPY AVAILABLE

INVESTIGATION 13: ARRHENIUS SOLUTIONS

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Name_____

Section_____

READING 13a: QUESTIONS

1. Order the acids and bases from Table 4 and 5 according to strength.

ACIDS BASES

3. Write the dissociation equations and the expressions for K_a and K_b for an acid and a base.

3. Predict which is a stronger acid, perchloric acid ($HClO_4$) or periodic acid (HIO_4). Explain the reasoning used in your prediction.

4. Predict which acid is stronger, carbonic acid (H_2CO_3) or boric acid (H_3BO_3). Explain the reasoning used in your prediction.

196

Acceptable Answers

1. The following listings are from strong to weak.

HCl—HNO_3—H_2SO_4
H_3PO_4—H_3AsO_4
HNO_2—$HC_2H_3O_2$
H_2CO_3—H_3BO_3
HCN—KOH—$NaOH$—$Ca(OH)_2$
$Zn(OH)_2$—$Pb(OH)_2$
$AgOH$—NH_4OH—NH_2OH

2. $K_A = \dfrac{[H^+] \times [HCO_3^-]}{[H_2CO_3]}$

$K_B = \dfrac{[K^+] \times [OH^-]}{[KOH]}$

3. $HClO_4$. Decreasing acidity is linked to an increase in the size of the central atom. Chlorine has a smaller atomic size than iodine; therefore, $HClO_4$ is a stronger acid than $HClO_4$.

4. H_2CO_3. Carbonic acid should be stronger than boric acid. Since carbon is a smaller atom than boron, and the smaller the central atom, the stronger the atom.

The central atoms in these acids take on a positive character when they are bonded in this manner with oxygen. Since repulsion of like charges increases over shorter distances, hydrogen is attracted less to the smaller central atom. The higher negative charge of the borate ion may also have an effect on the comparative acid strength.

BEST COPY AVAILABLE

INVESTIGATION 14: pH * ACIDS AND BASES

TEACHER'S GUIDE

INTRODUCTION

The students will explore various acid and base solutions to determine their hydrogen ion concentrations. This investigation will lead to the conceptual invention that the molar concentration of the H^+ ion determines the extent to which a solution is acidic or basic. This is measured in units called pH. If the pH is less than 7, the solution is acidic; if the pH is above 7, the solution is basic; and if the pH is equal to 7, the solution is neutral. READING 14: AN INTRODUCTION TO LeCHATELIER'S PRINCIPLE gives more information on dynamic equilibrium. The goal of the class discussion is to connect solution equilibria and strengths of acids as they are related to dynamic equilibrium.

CONCEPT

The molar concentration of the H^+ ion determines the extent to which a solution is acidic or basic and is measured in units called pH.

EQUIPMENT

A. Exploration

NOTE: Considerable preparation time is required for this investigation.

1. Labware

   a. 50ml beaker
   b. Dropper bottles for indicators
   c. 1000ml Erlenmeyer flasks for stock solutions
   d. 10ml pipettes with suction bulbs

2. Reagents

   Acid and base solutions for Exploration: The directions given here are for the preparation of 1 liter of stock solution. These stock solutions are to be diluted 10:1 with distilled water prior to their use by the students. Such dilutions permit a range of pH indications with the indicators used.

INVESTIGATION 14: pH * ACIDS AND BASES

Stock Solutions

a. Sulfuric acid--277ml concentrated H_2SO_4 per liter.
b. Acetic acid--5.73ml glacial acetic acid per liter.
c. Oxalic acid--63g solid per liter of solution.
d. Boric acid (H_3BO_3)--21g solid per liter of solution.
e. Sodium hydroxide--40g solid per liter of solution.
f. Potassium hydroxide--56g solid per liter of solution.
g. Calcium hydroxide--2 small chips of Ca metal reacted with 1 liter of water.
h. Citric acid--70g of solid per liter of solution.
i. Sodium carbonate (Na_2CO_3)--62g $Na_2CO_3 * H_2O$ per liter of solution.
j. Magnesium hydroxide ($Mg[OH]_2$)--0.1g MgO per liter of solution.

Indicators for Exploration: After the indicator solutions have been prepared, they should be put in small dropper bottles. One labeled set of indicator bottles should then be provided for each laboratory table.

a. Crystal violet--4ml of 1% solution in 100ml of distilled water.
b. Orange IV--0.1g in 100ml of distilled water.
c. Methyl orange--0.1g in 100ml of distilled water.
d. Methyl red--0.2g in 60ml of ethanol and 40ml of distilled water.
e. Bromothymol blue--prepared solution.
f. Phenolphthalein--0.5g in 50ml of ethanol and 50ml of distilled water.
g. Alizarin yellow R--0.1g in 100ml of distilled water.
h. Azo-violet--1g in 100ml distilled water.

INVESTIGATION 14: pH * ACIDS AND BASES

PRE-LAB INSTRUCTIONS

The colors indicated on the table should be interpreted rather loosely. Tell the students that the red indications are not always a deep red; in fact, they may appear orange sometimes. Further student experimentation should answer any questions regarding a particular indicator.

INVESTIGATION 14: pH * ACIDS AND BASES

Name_____

Section_____

INTRODUCTION

A scale that is used to describe the extent to which a solution is acidic or basic is called the pH scale. Numerical values on the pH scale range from 0 to 14. Certain substances called INDICATORS will show a given color inside the range of pH. Your task is to determine the range of the pH scale that corresponds to bases and the range of the pH scale that corresponds to acids, using indicators in base and acid solutions.

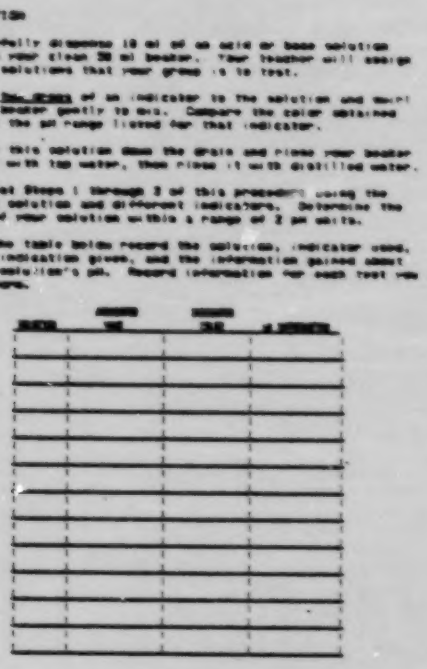
Nine indicators are listed below along with the colors they exhibit in ranges of pH.

Indicator	pH Range	Color
1.	1 to 2	yellow
	2 to 14	red
2.	1 to 13	red
	13 to 14	yellow
3.	1 to 4	orange
	4 to 14	yellow
4.	1 to 5	red
	5 to 14	yellow
5.	1 to 6	yellow
	6 to 14	red
6.	1 to 8	colorless
	8 to 14	red
7.	1 to 10	yellow
	10 to 14	red
8.	1 to 11	yellow
	11 to 14	violet

BEST COPY AVAILABLE

INVESTIGATION 14: pH * ACIDS AND BASES

EXPLORATION

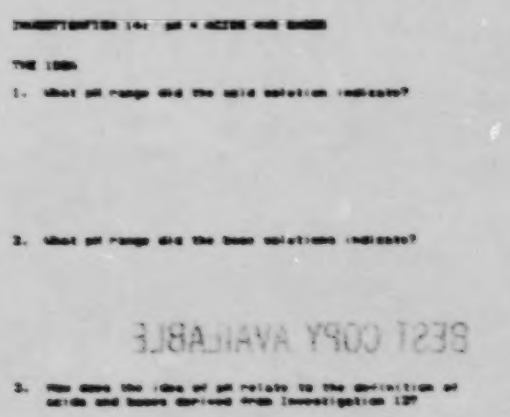


1. Point out the location of the solutions and indicators. Give instructions on the method to obtain the solutions. Assign the solutions to be tested by each group. If time permits, encourage the students to test other acid and base solutions.

Caution the students to be careful not to contaminate the indicator solutions.

On the chalk board or an overhead transparency, make a class chart for collection of data.

CONCEPTUAL INVENTION (THE IDEA)



Acceptable Answers

1. The acid solutions' pH range was from 0-6.

2. The base solutions' pH range was from 7.6 to 14.

3. An acid solution contains H^+ ions and indicates a pH between 0-6. A base solution contains OH^- ions and indicates a pH between 7.6 to 14.

BEST COPY AVAILABLE

INVESTIGATION 14: pH * ACIDS AND BASES

INVENTION DISCUSSION

A. List the acids and bases on the board. Have the class write down the ranges obtained from the tests. After the class has made a table of the ranges, ask them to analyze these data and formulate a statement about the pH of acid and base solutions. When the statement is made by a student and the class agrees to it, write the statement on the board.

B. The teacher should center the discussion on the solutions by asking the class how the idea of pH relates to their definition of acids and bases derived from laboratory work (see THE IDEA, Question 3). This will require time for the students to review their statements and correlate them with the questions being asked. Some guidelines for the direction of the ensuing discussion follows:

1. Hydrogen ions are released in acid solutions and hydroxide ions are released in base solutions.

2. pH involves a measurement of the concentration (or amount of) hydrogen ions that are in the solution. This will be a teacher invention.

3. The students' data will show the range of acid and base pH quite well, so the relationship of low pH with high hydrogen ion concentration can be made.

C. This is a good time to give an equation description of pH: the negative of the exponent of the hydrogen ion concentration expressed in moles per liter of solution (for example, a pH of 1 is derived from 1×10^{-1} moles of H^+ ion in a liter of solution). Molarity was introduced in EXPANSION OF THE IDEA, Investigation 12.

1. Ask the students what ion is characteristically predominant in base solutions. Lead the class to the idea that a base solution could contain hydrogen ions, but they are far outnumbered by hydroxide ions. Refer the students to the fact that base solutions have high pH values.

2. When the class has had time to accommodate the ideas above, ask them about the source of hydrogen ions in a base solution. This discussion will ultimately lead to the notion that water partially breaks up into hydrogen and hydroxide ions.

3. Based on the discussion to this point, the class should be ready to discuss the implications of a pH value of 7. The teacher should question the class about the relative hydrogen and hydroxide ion concentrations of such a solution. Discussion of this question should lead to the idea that pure water would have a pH of 7.

NOTE: The laboratory distilled water may not have a pH of 7. Even though we assume that the distilled water is pure--and it <u>is</u> for our practical purposes--there may be enough contaminants in it to cause some deviation from a pH of 7. This point may lead to some discussion about <u>purity</u> in the practical (or realistic) sense of the word.

4. Ask the students to compare the H^+ and OH^- concentration in water. When they conclude that they are equal, ask them to supply the numerical value which corresponds to this concentration (1×10^{-7} moles/liter).

5. Have the students multiply the hydrogen ion concentration by the hydroxide ion concentration. Invent the term--ionization constant--for water as 1×10^{-14}. Lead the class through the implications of this inverse relationship. (We assume that the pH of a solution added to the pOH of a solution equals 14. In other words, H^+ concentration multiplied by OH^- for most solutions would equal the value of the ion product for water. In this way one can determine OH^- concentration of a solution, if the H^+ concentration is known, and vice versa.)

INVESTIGATION 14: pH * ACIDS AND BASES

EXPANSION OF THE IDEA

Assign READING 14: AN INTRODUCTION TO LeCHATELIER'S PRINCIPLE, and Questions.

INVESTIGATION 14: pH * ACIDS AND BASES

Name_____

Section_____

READING 14

QUESTIONS

1. Responding to LeChatelier's Principle, explain what would happen to the equilibrium in each of the following systems, when the stated changes are imposed on the systems.

   a. $Pb(NO_3)_2 + NaCl \rightarrow PbCl_2 + NaNO_3$, when more NaCl is added to the system.

   b. $N_2 + 3H_2 \rightarrow 2NH_3$ + heat (all are gases), when heat is added to the system.

   c. A saturated NaCl in water solution, when half the solution water is evaporated.

2. Referring to Reading 5 in Investigation 13: THOSE MIGHTY ACIDS, why would a larger value for the equilibrium constant indicate that a reaction would go further toward completion than a reaction with a smaller value?

204

Acceptable Answers

1. (a) Assuming that the system is at equilibrium, adding more NaCl would shift the reaction to the right, producing more $PbCl_2$ precipitate.

(b) Assuming constant pressure, the reaction would shift to the left, thus relieving the stress on the right hand side. If pressure increased when the temperature increased, the shift in equilibrium would depend on whether the increased temperature or the pressure had a greater effect on the system. If the temperature effect is greater, the system would go to the left, but if pressure had the greater effect, the equilibrium would shift to the right (less molecules causing smaller pressure).

(c) The system would shift to the left of a reaction, such as NaCl [s] + $H_2O \rightarrow$ NaCl [aq].

2. Since the equilibrium constant is calculated by the concentration of products being divided by the concentration of the reactants, the further the reaction goes to to completion, the larger the equilibrium constant will be.

BEST COPY AVAILABLE

INVESTIGATION 14: pH * ACIDS AND BASES

A discussion of READING 14 should be held. The information presented here is intended to serve as a guideline for a class discussion. The goal of the class discussion is to connect solution equilibria and strengths of acids as they are related to dynamic equilibrium.

1. Ask the students what the ionization constant for water is. (1×10^{-14})

2. Since the ionization constant for water is 1×10^{-14}, what values do we use for $[H^+]$ and $[OH^-]$? (1×10^{-7} for each)

3. Explain, in terms of LeChatelier's Principle, what happens to this equilibrium when an acid dissolves in water.

$$HOH \longrightarrow H^+ + OH^-$$
$$\longleftarrow$$

When an acidic solution is added, the equilibrium is stressed forcing recombination of water molecules. The resulting solution will have an excess of H^+ ions.

4. Explain, in terms of LeChatelier's Principle, what happens to this equilibrium when a base dissolves in water.

$$HOH \longrightarrow H^+ + OH^-$$
$$\longleftarrow$$

When a basic solution is added, the equilibrium is stressed forcing recombination of water molecules. The resulting solution will have an excess of OH^- ions.

5. Given: A solution with a pH value of 2. If a base is added to this system, what happens to the equilibrium?

When a basic solution is added, the OH^- ions will combine with the H^+ ions forming water molecules which will cause the equilibrium to shift toward the left.

6. Given: A solution with a pH value of 6. If an acid is added to this system, what happens to the equilibrium?

When an acidic solution is added, the H^+ ions will combine with OH^- ions in solution resulting in a shift of the equilibrium to the left.

7. If LeChatelier's Principle is applied to the following system

   $N_2 + 3H_2 \rightarrow 2NH_3 + heat$,

   what effect on the system will each of the following conditions have?

   a. Increasing one of the reactants (shift equilibrium to the right)
   b. Decreasing the product, NH_3 (shift equilibrium to the right)
   c. Heating the reactants (shift equilibrium to the right)

INVESTIGATION 15: ACID-BASE INTERACTION

TEACHER'S GUIDE

INTRODUCTION

This Investigation is designed to introduce the ideas of acid-base neutralization. There are two possible explanations for acid-base neutralization: a physical mixing or a chemical interaction. The interaction in this investigation will be discussed from both points of view and suggestions of ways to distinguish between them will be made. The use of concentration to determine the equivalent number of moles is a point to be emphasized in the discussion. Discuss the arrival of the neutralization point which is called the end point or equivalent point.

CONCEPT

Acids and bases can interact with each other to produce a neutral solution.

EQUIPMENT

A. Exploration

1. Labware

a. Beakers--50 ml and 100 ml
b. Dropper pipet
c. 2-1,000 ml Erlenmeyer flasks
d. 2-10 ml pipettes with suction bulbs

2. Reagents

a. Acetic acid (1 liter)--5.7 ml of glacial acetic acid per liter of solution
b. Sodium hydroxide--8.0 g of solid NaOH per liter of solution

B. EXPANSION OF THE IDEA

Demonstration

1. Labware

a. Magnetic stirrer with magnet bar
b. 1000 ml beaker
c. Light bulb conductivity apparatus
d. Ring stand

2. Reagents

a. Sulfuric acid solution (750 ml)

This solution is prepared by adding 10 ml of the stock solution for Investigation 14 to 990 ml of distilled water to make a liter of solution.

b. Barium hydroxide solution (150 ml)

Heat a liter of distilled water to boiling and then transfer the water to a magnetic stirrer. Begin stirring and add at least 60 g of $Ba(OH)_2 \cdot H_2O$. Cover the container and allow the stirrer to run for 15 minutes. Turn the stirrer off an allow the precipitate to settle for a few minutes. Filter (through double paper) into a container that can be tightly sealed.

c. Phenolphthalein indicator solution

d. Bromothymol blue indicator solution

1. Put 10 ml of acetic acid solution in a clean 100 ml beaker. Add four drops of methyl orange indicator solution and swirl to mix.
2. What is the color of the resulting solution?_____
3. In what range of pH does the solution lie?_____
4. Add four drops of phenolphthalein indicator solution to the mixture above and swirl to mix.
5. Does the addition of phenolphthalein have an effect on the color of this solution? Explain your response.
6. Put 7 ml of sodium hydroxide solution into a clean 50 ml beaker.
7. Using a clean dropper pipet (rinse with distilled water), slowly add three drops of sodium hydroxide solution to the acid solution and swirl to mix. Record your observations.
8. Continue adding sodium hydroxide solution (dropwise with swirling) until you have added all the sodium hydroxide solution. Write a detailed summary of your observations during this procedure. (NOTE: You may be wise to test the solution with litmus paper from time to time.)

CONCEPTUAL INVENTION (THE IDEA)

INVESTIGATION 15: pH + ACIDS AND BASES

THE IDEA

1. Did this procedure involve a chemical change between the acid and base? Explain your response thoroughly. If so, write a balanced equation for the reaction.

2. What does the beaker contain after you have added three drops of NaOH solution to the acetic acid?

3. What does the beaker contain when you have added enough NaOH to cause the color to fade from orange to yellow?

4. What does the beaker contain when you have added enough NaOH to cause the color to change to a dark pink?

Acceptable Answers

2. Orange
3. 0 to 4.4
5. No, phenolphthalein is colorless below a pH of 10.

7. The solution turns red when the drop goes in, but upon swirling the solution, it goes back to orange.

8. Solution gradually turns yellow. With one final drop, it turns red.

Acceptable Answers

1. Yes, the indicators show us that new substances were formed; or, no because the OH^- ion concentration overpowered the H^+ ion concentration.

$HC_2H_3O_2 + NaOH \rightarrow NaC_2H_3O_2 + HOH$

2. Indicators, $HC_2H_3O_2$, NaOH (if physical),
H_2O, $NaC_2H_3O_2$ (if chemical)

3. Same as in Number 2

4. If chemical is assumed, NaOH, $NaC_2H_3O_2$, H_2O, and indicators with NaOH in excess.

BEST COPY AVAILABLE

INVESTIGATION 15: ACID-BASE INTERACTION

INVENTION DISCUSSION

A. Through a class discussion obtain the observations made during the laboratory exercise.

B. Question the students about the reactants, $HC_2H_3O_2$ and NaOH. Review the definitions for Arrhenius acids and bases.

C. Ask the students if acetic acid and sodium hydroxide reacted chemically. If so, what type of chemical reaction was it?

D. Since the interaction of acetic acid and sodium hydroxide is a double displacement reaction, the products should be easily named by your students. Discuss the use of the indicators in solutions and the significance of their indications. Compare the pH of the reactants to the pH of the products. Invent the term <u>acid-base neutralization.</u> The invention is that acids and bases can interact with each other to produce a neutral solution.

E. Write the equation for the chemical equation on the board.

$HC_2H_3O_2 + NaOH \rightarrow HOH + NaC_2H_3O_2$

Ask a student to read this equation in terms of molecules, and then in terms of moles. Discuss the number of moles of acid and base and the use of concentration to determine equivalent number of moles.

F. Discuss the arrival of the neutralization point and invent the term end point or equivalent point.

G. Discuss and list the characteristics of acid-base neutralization reactions.

1. Arrhenius acids and bases
2. One product is always water
3. Double displacement reaction
4. The product other than water is a salt
5. The resulting solution should be neutral

INVESTIGATION 15: ACID-BASE INTERACTION

EXPANSION OF THE IDEA: BARIUM HYDROXIDE SULFURIC ACID

INTRODUCTION

One of the benefits of this exercise is to demonstrate conclusively that an acid-base interaction is a chemical change. Is this interaction physical or chemical? The students should be completing Student Page 208 prior to and during the demonstration.

DEMONSTRATION

NOTE: It is strongly suggested that this demonstration be practiced prior to performance in class.

1. Place the beaker on the magnetic stirrer and place the magnet bar in the beaker.

2. Pour about 250 ml of sulfuric acid solution into the beaker and then turn on the stirrer and adjust the speed as desired.

3. Support the conductivity apparatus on a ring stand and then lower the electrodes into the acid solution. Plug in the apparatus.

4. Pour two or three ml of phenolphthalein indicator solution into the beaker. Add an equal amount of bromothymol blue indicator solution.

5. Begin by slowly adding barium hydroxide solution to the beaker.

6. When the endpoint has been passed and the light is burning brightly in the basic solution, the teacher may wish to reverse the process by slowly adding more acid solution to the beaker.

NOTE: The barium hydroxide solution should be kept covered when not in use. The solution tends to react with carbon dioxide from the air to form insoluble barium carbonate.

INVESTIGATION 15: ACID-BASE INTERACTION

EXPANSION OF THE IDEA: BARIUM HYDROXIDE + SULFURIC ACID

DEMONSTRATION

1. Diagram and label the experimental set-up used in this demonstration.

2. Describe what happens as barium hydroxide—$Ba(OH)_2$—is added to the sulfuric acid solution—H_2SO_4.

QUESTIONS

1. What evidence shows that this interaction is a chemical reaction?

2. Write a balanced chemical equation for the reaction.

3. At what point in this demonstration is there an equivalent amount of acid and base in the beaker? What is the pH of the solution at this point?

Allow the students time to complete Step 1 prior to beginning the demonstration.

Step 2 should be filled in during the demonstration.

Assign the questions immediately following the demonstration.

Acceptable Answers

1. Formation of a precipitate, change in temperature, change in pH, change in conductivity.

2.

$$H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2H_2O$$

3. The pH at this point is below 7.6 because the solution was still yellow.

DISCUSSION OF DEMONSTRATION

1. This demonstration clearly shows that the interaction of acids and bases is a chemical reaction. One of the major advantages of practicing this demonstration beforehand is that the teacher can plan the timing of the discussion items that are integral parts of the demonstration.

2. Indications of chemical reaction in this demonstration include:

   a. Variation in intensity of the light from the conductivity apparatus.
   b. Formation of a precipitate.
   c. Change in pH, as indicated by the phenolphthalein and bromothymol blue color change. (Litmus may also be used.)
   d. Change in temperature.

   During the demonstration, allow the students to identify these occurrences as evidences to support the idea of a chemical reaction.

BEST COPY AVAILABLE

3. The teacher should ask the students to suggest products of the reaction and to explain how their observations of the demonstration would support their choice of products. Write these on the board.

4. The teacher should review the terms acid-base neutralization and salt.

   a. Ask for several examples of salts that students have encountered. Ask if aluminum chloride and sodium fluoride are salts. (Yes.)

   b. Ask how these salts are different from the majority that are mentioned by the students. ($AlCl_3$ forms an acid solution and NaF forms a base solution with water. Most of the other salts form neutral solutions.)

   c. Invent the terms acid-forming salt and base-forming salt.

5. Assign the Reading, DOES AN ACID-BASE NEUTRALIZATION RESULT IN A NEUTRAL SOLUTION?

INVESTIGATION 16: SALTS AND WATER

TEACHER'S GUIDE

INTRODUCTION

This investigation contains activities which lead to the invention of the Bronsted-Lowry Acid-Base Theory. A Bronsted-Lowry acid is a proton donor and a base is a proton acceptor. This is the most sophisticated acid-base theory which will be developed through investigation. If the teacher wants to discuss the Lewis Acid-base Theory, the appropriate time would be after Investigation 17.

CONCEPT

Compounds that donate H^+ ions in a reaction are acids, and compounds that accept H^+ ions in a reaction are bases, according to definitions advanced by Bronsted and Lowry.

EQUIPMENT

A. Exploration

1. Labware

   a. Plastic spoons
   b. 50 ml beakers
   c. Red and blue litmus paper
   d. Light bulb conductivity apparatus

2. Reagents

   a. Sodium carbonate
   b. Ammonium chloride
   c. Sodium acetate
   d. Sodium phosphate
   e. Ammonium acetate
   f. Distilled water

B. Expansion of the Idea: Demonstration

1. Labware

   a. pH paper
   b. 250 ml beakers--8

2. Reagents

a. 0.5M Na_2HPO_4
b. 0.5M NaH_2PO_4
c. 0.1M H_2SO_4
d. 0.2M NaOH

EXPLORATION

[illegible]

1. [illegible]

Salt	Acid (name and formula)	Base (name and formula)
[illegible]		
[illegible]		
[illegible]		
[illegible]		
[illegible]		

[illegible]

[illegible]

2. [illegible]

Salt	Litmus Test	Conductivity Test

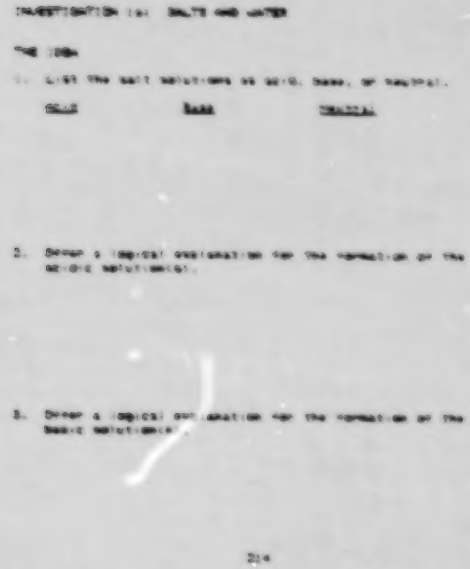
[illegible]

1. The only necessary preparation is to put out the salts that the students are to test.

BEST COPY AVAILABLE

INVESTIGATION 16: SALTS AND WATER

CONCEPTUAL INVENTION (THE IDEA)



Acceptable Answers

1.

acid	base	neutral
NH_4Cl	$NaC_2H_3O_2$	
$NH_4C_2H_3O_2$	Na_2CO_3	
	Na_3PO_4	

2. To be acidic the solution must have an excess of H^+ ions. The positive ion of the salt must attract the OH^- ion.

3. To be basic the solution must have an excess of OH^- ions. The negative ion of the salt must attract the H^+ ion.

INVENTION DISCUSSION

A. Begin the discussion by reviewing the data concerning the acid and base solutions.

B. Review the defining criteria for acid and base solutions (for example, excess of H^+ and OH^- ions, respectively).

C. Lead the class in an analysis of the compound formulas involved to determine the source of excess H^+ or OH^- ion in the solutions.

NOTE: It will be useful during the discussion to define the hydronium ion and talk about its relationship to the H^+ ion discussed before.

BEST COPY AVAILABLE

For example:

1. What were the results of your tests when the salt, Na_2CO_3, was added to water? (A basic solution formed.)

2. If Na_2CO_3 and water form a basic solution, what must be in the solution? (An excess of OH^- ions.)

3. Write in equation form how this would happen.

   $Na_2CO_3 + HOH \rightarrow 2Na^+ + HCO_3^- + OH^-$

4. How does HCO_3^- form?

   $CO_3^{2-} + H^+ \rightarrow HCO_3^-$

5. Na_2CO_3 is a base-forming salt. A base-forming salt is a salt whose water solution has an excess of OH^-.

6. A general rule to follow is:

   a. If the charge on the positive ion is a +3, an acidic solution will form. If the charge on the positive ion is a +2, a neutral solution will form. If the charge on the positive ion is a +1, either a neutral or a basic solution will form.

   b. If the charge on the negative ion is -2 or -3, a basic solution will form.

D. When both acid and base solutions have been explained through discussion, invent the Bronsted-Lowry acid-base definitions. The reading, ALL DONATIONS ARE ACCEPTED, offers a thorough discussion of Bronsted-Lowry acid an base definitions.

E. Point out the necessity for the presence of a Bronsted-Lowry acid and base in a reaction.

EXPANSION OF THE IDEA

The purpose of the demonstration with the phosphatse compounds is to show some more examples of the Bronsted-Lowry acid-base theory. Dynamic equilibrium ideas and Le Chatelier's Principle can also be brought into this discussion.

INVESTIGATION 16: SALTS AND WATER

NOTE: Have 0.5M Na_2HPO_4 and NaH_2PO_4 solutions made, as well as 0.1M H_2SO_4 solution and 0.2M NaOH solution.

1. Review with the students the meaning of the concentrations of the several solutions. Especially point out the reason for the acid and base concentration difference.

2. Test each of the four solutions with pH paper and get consensus on the pH value of each. Encourage students to note these changes.

3. Engage in a discussion of the interaction of NaH_2PO_4 with water to produce the pH values obtained for each solution. Attempt to bring in the Le Chatelier principle in discussing how each solution's pH differs from the pH of water.

4. Carry on a discussion similar to the one above after mixing the two solutions. Discuss the interactions involved to arrive at the pH of the mixed solutions.

5. Carry out the 8 tests listed on the student note sheet, and then engage in a discussion of the results of those test. Use the Bronsted-Lowry terminology, as appropriate, in the discussion.

The students should complete these sheets during and after the demonstration.

1. $Na_2HPO_4 + HOH \rightleftharpoons H_2PO_4^- + 2Na^+ + OH^-$

$NaH_2PO_4 + HOH \rightleftharpoons Na^+ + HPO_4^{2-} + H_3O^+$

BEST COPY AVAILABLE

INVESTIGATION 16: SALTS AND WATER

Assign READING 16: ALL DONATIONS ARE ACCEPTED, and its questions.

[illegible]

1. In investigation 5 you added water to magnesium nitride, which had formed in your system, and got a reaction that can be described by the following equation:

   [illegible]

   a. List the spectator ion(s) in this reaction.

   b. How many single proton transfer reactions would be involved in making this reaction occur?

   c. Write a balanced equation for each of the single proton transfer equations. (NOTE: The sum of these equations must equal the overall balanced equation above.)

2. Ammonium carbonate will decompose at atmospheric pressure to produce ammonia gas (NH_3). Write a balanced equation for this decomposition. Label the reactant acid and base and the conjugate acid and base.

3. Label as Brønsted-Lowry acids and bases the appropriate parts of the equation showing the reaction of barium hydroxide with sulfuric acid.

Acceptable Answers

1. a. Mg^{2+}
   b. 6 proton transfers
   c. $N^{3-} + H_2O \rightarrow NH^{2-} + OH^-$
   $NH_2^- + H_2O \rightarrow NH_3 + OH^-$
   $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

NOTE: This set of equations can be repeated for the other nitride ion, or each of the equations above can be doubled.

2. $(NH_4)_2CO_3 \rightarrow H_2CO_3 + 2NH_3$
   acid base acid base

3. $Ba(OH)_2 + H_2SO_4 \rightarrow BaSO_4 + 2H_2O$
   base acid base acid

Notice that the hydroxide ion is the base on the left-hand side of the equation and the sulfate ion is the base on the right-hand side. The barium ion is a spectator.

BEST COPY AVAILABLE

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

TEACHER'S GUIDE

INTRODUCTION

There are two main objectives in this series of discussions: (1) Introduction of the concept of covalent bonding; and (2) Rationalization of molecular shapes for covalently bonded molecules.

In order to achieve these main objectives, several conceptual inventions and expansions of ideas will be employed (see Content Summary). As a result of the learning in these discussions, the investigation of certain physical properties can be done in terms of the molecular structure of covalently bonded substances.

As with the ATOMIC STRUCTURE discussions (Investigation 6), this series of discussions is divided by concepts introduced rather than by time blocks. The teacher's concern, therefore, should be more with the concept understanding gained by the students than with the amount of time required to complete the discussions. Furthermore, this series of discussions also departs from the usual format for the investigations. There is one exploration activity at the beginning, however, and there are several worksheets that go along with the discussions. The teacher should continually strive to maximize student input into these discussions.

CONCEPT

Covalent bonding is a chemical bonding created by the mutual attraction for electrons between atoms in an element or compound.

CONTENT SUMMARY

A. CHARACTERISTICS OF METHANOL (EXPLORATION)

B. First discussion

   1. Covalent bonding

C. Second discussion

   1. Electron pairing
   2. Electron spin
   3. Electron energy sublevels (s, p, d, f)

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

D. Third discussion

1. Electron orbital
2. Bonding orbital
3. Electron-dot structures (Lewis)

E. Fourth discussion

1. Single, double, triple covalent bonds
2. Electron affinity
3. Electronegativity

F. Fifth discussion

1. Polarity
2. Non-polarity

G. Sixth discussion

1. Permanent dipoles
2. Linear molecule
3. Planar molecule
4. Tetrahedral shape

H. Seventh discussion

1. Shapes of various molecules

EQUIPMENT

A. For CHARACTERISTICS OF METHANOL

1. Methanol--1 liter
2. Distilled water--in wash bottles
3. Light bulb conductivity apparatus--1 per table
4. Ring stand--1 per table
5. 250ml beakers--15
6. 150ml beakers--15
7. Litmus paper--centrally supplied

B. For Demonstrations During Discussions

1. 250ml wash bottles--3
2. 400ml beakers--3
3. Glass and rubber rods--1 each
4. Silk and wool cloth--1 each
5. Round balloons--5
6. Disc magnets--4
7. Methanol--100ml
8. Carbon tetrachloride--100ml
9. Distilled water--100ml

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

INTRODUCTION

During INVESTIGATION 13: ARRHENIUS SOLUTIONS, you developed ideas that allowed you to recognize and specify whether a substance was an acid or a base. In this investigation you will be studying one compound: methanol (CH_3OH). In Steps 1 and 2 you will predict some characteristics based on your previous study, and then you will test those predictions in the laboratory in Steps 3 and 4.

EXPLORATION

1. State how methanol will affect litmus paper and whether or not it will conduct electricity. Give supporting evidence for your predictions.

2. State how a solution of water and methanol will affect litmus paper and whether or not it will conduct electricity. Will the concentration of solution affect those properties? Give supporting evidence for your prediction.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

3. Test methanol for conductivity and acid-base characteristics. Describe your tests and record the results of your tests.

4. Test solutions of methanol and water for conductivity and acid-base characteristics. Describe your tests (including the relative proportions of the two liquids) and record the results of your tests.

THE IDEA

1. Did one or more of the solutions of methanol and water ionize? If so, what ions were formed? If not, why not?

2. Offer an explanation for the results of the tests in light of the fact that methanol is soluble in water.

BEST COPY AVAILABLE

One period should be sufficient time for the students to take the data and answer the questions under THE IDEA.

The response to the two sets of predictions will be diverse. Most people, however, will probably relate the formula for methanol to the characteristics of acids and bases.

The predictions should be completed prior to the students being allowed to perform the laboratory procedures.

The teacher should dispense the methanol samples to the students. The sample of approximately 50ml should be an ample supply for the students to use in all their tests, if they plan their tests ahead of time.

CONCEPTUAL INVENTION (THE IDEA)

Satisfactory responses to the two questions will be given here, but the conceptual invention will be covered in the discussion notes that follow.

Question 1: None of the methanol and water solutions contained enough ions to show conductivity. We can, therefore, assume that the methanol did not ionize in water solution. There were no acid or base litmus tests. An acceptable reason that the ionization did not occur would be that the water did not have a strong enough attraction to pull the methanol molecules apart into ions.

Question 2: There are several ways that students could approach this question, and it is suggested that the teacher accept any reasonable attempt at answering it. Some students may suggest that, since methanol contains the same elements as sucrose, the water solution did not conduct because the two compounds have similar soubility characteristics.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

GUIDE TO THE COVALENT BONDING DISCUSSIONS

I. COVALENT BONDING: FIRST DISCUSSION

A. Ask the class why the solution did not conduct electricity. Focus their attention, if necessary, on the lack of ion formation in the solution process. (They have observed this once before, when they dissolved sugar in water. Solutions of boric acid did not conduct either, but they did indicate acid to litmus paper.)

B. Since the formula for methanol indicates an OH, ask the class why this OH does not appear as a hydroxide ion in the water solution. (It is held too tightly by the rest of the molecule.)

NOTE: Sometimes a student will suggest the term covalent bonding to explain why the OH is held so tightly. This student can be questioned about what covalent bonding is and how it occurs. Such questioning may lead to the invention of the concept for the rest of the class, but it may not be worthwhile to pursue the invention yet; the teacher will have to decide this issue. If no one suggest the term or the discussion of the term becomes fruitless, the following guidelines are suggested to work toward the conceptual invention.

C. Ask the class how many electrons a neutral carbon atom has (6).

1. Ask which of these electrons will be involved in chemical reactions (valence electrons).

2. Propose that since only the valence electrons would be involved in chemical reactions, we will represent only those electrons--as dots--around the symbol for the element. Do this on the board.

The structure is reasonable if we assume that the electrons will repel each other as much as possible.

3. Repeat this questioning for oxygen and hydrogen.

BEST COPY AVAILABLE

NOTE: Do not attempt to put the dots in pairs at this time. Simply arrange the dots around the symbols. The important thing in this discussion is to get to the concept of electrons being shared between atoms. Electron pairing will be discussed in the second discussion.

D. Since the students have been working with forming compounds by the attraction of oppositely-charged ions, ask the class what possible noble gas configuration ions could form for the elements in methanol (C^{4+}, C^{4-}, O^{6+}, O^{2-}, H^{+}, H^{-}). The H^{+} ion has no electrons, of course, but it is a common ion of hydrogen.

1. Ask the class to attempt to combine four hydrogen, one carbon, and one oxygen ion from this list to form a neutral methanol molecule. (There are a couple of ways that this can be done, if both types of hydrogen ion are assumed.)

2. When the suggestions are exhausted and noted on the board, work with the class to ascertain the feasibility of the different molecules.

3. One of the suggestions will have to contain the O^{6+} ion. Remembering how multiple ionizations require large amounts of energy, students will often suggest that the molecule with the O^{6+} ion is not reasonable.

4. The other molecule, which contains carbon with a positive four charge, does not seem so unreasonable to many students. Since ionization energy considerations have come up, the teacher can conveniently use them to discuss this second molecule.

   a. In this second molecule, three of the hydrogen ions will be negative and one will be positive. The oxygen ion will be negative.

b. Starting with a dot diagram for the neutral carbon atom, ask the students if it is reasonable that a hydrogen atom could pull an electron off the carbon atom. (When the students compare first ionization energies for the two elements, they will find that hydrogen has greater attraction for its electron than carbon does for the first electron to come off of it. Students generally conclude that such ionization--forming a negative hydrogen ion and a carbon ion with a one positive charge--is reasonable.)

c. Proceed by having another hydrogen atom pull off a second electron from the carbon. When the students compare the second ionization energy of carbon with the first ionization energy of hydrogen, they see immediately that this model will not work because now carbon is attracting its remaining valence electrons more strongly than the hydrogen is attracting its electron.

E. At this point, the teacher can propose an alternative method for building the molecule from neutral atoms, since the ion method will not work an since methanol does exist.

1. Begin with the dot structure for the neutral carbon atom. Draw one hydrogen atom near the carbon atom, such as H:C•.

2. Ask if hydrogen would attract the carbon electron that is shown near to it. Require support for the answer given. (Yes, hydrogen will attract carbon's electron. Since hydrogen has an ionization energy, it attracts its own electron. It would also, therefore, attract other electrons that were close to it.)

3. In a similar manner, ask if the carbon nucleus would attract hydrogen's electron. When agreement has been achieved on the mutual attraction of atoms for other atoms' electrons, continue adding atoms on the carbon atom until the molecule is complete.

4. Ask if this drawing indicates a neutral molecule. Get them to agree that it does before proceeding.

5. The problem now is to show this model for the molecule to be reasonable to the students in terms of their knowledge about atomic structure.

NOTE: The foregoing suggests that the hydrogen electron (dot) was drawn near one of the dots of carbon, when the teacher drew the model on the board. To help center the students' attention on the interaction discussed above, the teacher may want to circle this pair of dots on the diagram.

6. Once the idea of mutual attraction for the electrons is understood, the teacher, using class input, can show the other mutual attractions that hold the molecule together by circling the pairs of electrons that are mutually attracted.

7. Review the structure now by showing how each atom in the compound can be shown to have noble gas electron configurations because of the sharing of electrons. This sharing of electrons implies an overlapping of the valence levels of the atoms involved.

8. Invent the concept of covalent bonding. <u>Covalent bonding is a chemical bonding created by the mutual attraction for electrons between atoms in an element or compound.</u>

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

II. COVALENT BONDING: SECOND DISCUSSION

A. Much of the work that will be done in this discussion is didactic in nature. The teacher should be careful to rely, as much as possible, on the present understandings of the students. It is of equal importance that the teacher elicit as much response and feedback from the students as possible. This will help ensure that the students are following the reasoning used in the discussion. The teacher should use the chalkboard freely to write down important items during the discussion. This discussion will proceed through the introduction of the concepts of electron pairing and energy sublevels.

B. Briefly review the following points from the first discussion.

1. Covalent bonds are formed between atoms, when electrons from each atom are shared between the two atoms.

2. For representative elements, this sharing results (in general) in the formation of noble gas electron configurations for the atoms involved in bonding.

3. Covalent bonding occurs due to sharing of some or all of the valence electrons of the atoms involved.

C. Three questions that should be dealt with in the present discussion are: (1) Is it reasonable that electrons would form pairs? (2) How does the idea of electron pairing relate to the established atomic model? (3) Will electrons pair every time they get the chance?

NOTE: These questions should not be introduced all at once. They should be presented as they become applicable to the discussion. It would be good to write them on the board as they are discussed.

D. Question 1: The teacher may wish to refer to the diagram for methanol, where the electrons shared between atoms seemed to be unique to a given covalent bond. This representation implies that two electrons involved in the covalent bond existed relatively close to each other.

1. This can be treated by making some assumptions about the electron.

NOTE: This will have to be done by the teacher.

   a. Assume that the electron has an imaginary axis through it and that it spins on this axis. (The kinetic energy due to this spinning is part of the total energy of the electron.)

   b. A moving charge generates a magnetic field. We have found it helpful to demonstrate this idea by wrapping many turns of insulated wire around a large iron nail to make an electromagnet. After a brief review about current being the movement of charge, the ends of the wires are briefly connected to the poles of a large dry cell and the magnetism in the nail is demonstrated. (The magnetic field development would apply to any movement of the electron, but the students' attention should be focused on the spinning movement only.)

   c. This magnetic field, produced by the electron spinning, could further be assumed to produce a "north" and "south" magnetic pole.

2. Using these assumptions, the teacher can show how the electrons can have a magnetic attraction for each other (finally obtaining a diagram such as: ⇌). Two important points should be made:

   a. Even though there is magnetic attraction between two electrons, their <u>electrical charges are still the same and repulsion still exists between them.</u> The balance between the attractive magnetic forces and the repulsive electrical forces allows them to exist near each other.

   b. A pair of electrons will always have its spins aligned oppositely, hence the representation above.

3. Be sure that the class has begun to accommodate these ideas before going on the next question.

E. Question 2: As an aid to the understanding of this part of the discussion, the teacher should instruct the class to look at the graphs of first ionization energy that were plotted earlier.

1. The idea of electrons pairing in covalent bonds implies that the electrons in a neutral atom or a monatomic ion (for simplicity) could also be paired.

2. Ask the class if this idea of electrons pairing in an atom would have any adverse effect on their model for the atom. (No.)

NOTE: The teacher may want to use this opportunity to question the often presented model in which the electrons are perceived as moving in "circular orbits" around the nucleus. This point should be handled carefully because it constitutes a major revision to the model being used by many students. Furthermore, we will return to this consideration later in our discussions of bonding.

F. Question 3: Have the class examine the ionization energy graph for the second period elements.

1. Begin at lithium and move toward the right. (Diagrams on the board such as Li, Be, etc., might be useful in the early going.)

2. At boron, the ionization energy drops somewhat. Ask the class if the two valence electrons in beryllium could be paired. When the class agrees to this, ask them what is implied by the dip in ionization energy for boron.

NOTE: The point to be made here is an important one, but it is also a point that students may have difficulty receiving. The teacher should be ready to review the idea of ionization energy to insure that the students realize that the third valence electron of boron actually has more energy than the other two.

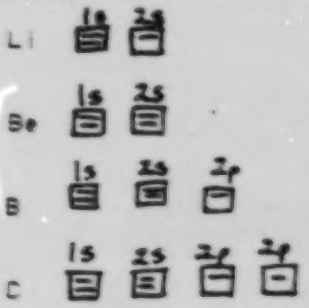
3. The teacher can now invent the concept of <u>energy sublevel</u> for energy levels that can contain more than two electrons. The point should be made that energy sublevels are part of a main energy level of electrons. (An analogy to a split-level house may be useful; that is, one house, but different levels within the house.)

4. It is reasonable to ask how many energy sublevels there are in the second energy level. Allow the class to puzzle over this question, using the ionization energy data, before proceeding to establish a model that represents two energy sublevels for the second energy level (for example, "Let's assume that there are two.").

5. Ask the class how many electrons could possibly occupy the higher energy sublevel, using the assumption that there are only two sublevels in the second energy level.

6. When the class has determined that six electrons could occupy the higher energy sublevel, lead them through the second period elements, using box diagrams for the electrons. (Save the invention of <u>orbitals</u> for the next discussion.)

Box diagrams:



7. The box diagram for carbon will undoubtedly cause a reaction from the class because they will want to pair the last two electrons. Show them that, according to that reasoning, they should expect to get a dip in ionization energy for nitrogen (as it occurred for boron since beryllium was paired), but that the dip in ionization energy does not occur until oxygen. Furthermore, if one assumed pairing of the last two electrons in carbon (and if the ionization energy data supported that pairing), we would be forced to assume that the second energy level contained more than two sublevels. Further support for not pairing the last two electrons is that the three spaces in the higher sublevel are equivalent. Since electrons repel, the least stressful position for the second electron would be in a separate space.

NOTE: There may be some concern about the significance of the dip in ionization energy at oxygen. Since the

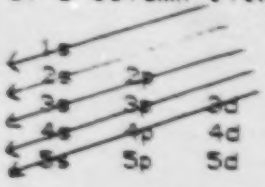
idea of pairing the first two electrons in the second sublevel does not work in terms of the ionization energy data, and since we are working with the assumption that electrons in an atom will pair, the dip at oxygen must be explained in a different way. Suggest to the students that a half-filled sublevel is somewhat harder to ionize than one that as one more electron (a sort of stability of half-filled sublevels).

8. Continue with the box diagrams through neon. Discuss the stability of filled sublevels in terms of the high ionization energy of neon and also in terms of the ionization energy of beryllium (as compared to that of boron).

9. Introduce the sublevel designations: "s" for sublevels that can contain a maximum of two electrons; "p" for sublevels that can contain a maximum of six electrons; "d" for sublevels that can contain a maximum of ten electrons; and "f" for sublevels that can contain a maximum of fourteen electrons. Write these designations on the box diagrams.

6. ASSIGNMENT: Have the students write box diagram electron arrangements for the first thirty-six elements (hydrogen through krypton). The boxes should be labeled with the energy level number and the sublevel designation.

NOTE: When these papers have been completed, it should become a matter of discussion about the order of filling of the sublevels when the B-column elements are reached. The students should be referred to the development of the filling order that was made during the ATOMIC STRUCTURE discussions. It would be appropriate to show the students the diagonal rule for sublevel filling after it is clear to them about the arrangement of B-column element electrons.



III. COVALENT BONDING: THIRD DISCUSSION

A. This discussion builds on the ideas developed in the first two discussions by the conceptual inventions of orbital and electron-dot structures for atoms and molecules. The teacher may wish to collect the box diagram exercises from the students, but it would be useful to review the exercises briefly to insure that the students have been able to work with the ideas discussed in the second discussion.

B. In the course of the brief review mentioned above, the teacher should write some of the box diagrams on the board.

1. Point out that the boxes are not meant to represent the actual shape of the arrangement of the electrons around an atom. They are there to show how electrons have the capability of pairing within an energy level or sublevel. This is an important point because, as the atomic model is modified, the students need to be reminded of the applicability of revisions to their existing ideas about the model.

2. Ask the class to explain how one pair of electrons in an atom would interact with the other pairs of electrons (or single electrons) in the same atom. Lead the class, through discussion, to the idea that these pairs of electrons (or single electrons) would repel each other as much as possible--that they would move to positions that are farthest away from the other electrons. Remind the students of the larger effective space occupied by the electrons due to their charges (ref: ATOMIC STRUCTURE discussion).

3. Invent orbital (write it on the board) as the effective space occupied by a pair of electrons--or one electron, if it is not paired in the atom.

4. Specify that the boxes represent electron orbitals and have the class state the number of s, p, and d orbitals there are in the first, second and third energy levels. (One s orbital in the first, one s orbital and three p orbitals in the second, and one s orbital, three p orbitals, and five d orbitals in the third.)

5. Continue this questioning by asking how many electrons could occupy these orbitals. (A maximum of two electrons in each orbital.)

C. Direct the attention of the students to the box diagram for neon, showing that all orbitals contain a pair of electrons. Point out that all the valence orbitals in neon are filled and ask them to state a relationship between this observation and the chemical characteristics of neon.

1. Draw a box diagram for helium and ask the class if the chemical characteristics of helium can by explained in the same way that those characteristics were explained for neon.

2. Draw a box diagram for beryllium and ask the class why beryllium does not behave chemically like the noble gases. (Do not show the p orbitals when asking this question.)

NOTE: This is an important point for the students to understand. An astute student may provide the answer quickly, but the teacher should carefully follow up on such a response to insure that the entire class grasps the idea. The availability of p orbitals in the second, and higher, energy levels will become important in the rationalization of the electron dot structures.

D. Question: "Could box diagrams be used to explain the chemical combination of hydrogen with chlorine to form hydrogen chloride?"

1. Have the class write box diagrams for these two elements on their note paper (and the teacher can then write them on the board from instructions given by the students).

2. Since the students know that this compound will form, ask them which electrons from the two atoms can be involved in the chemical combination (circle or bracket the valence electrons on the board diagram as they respond).

3. Ask them if it is possible for this compound to form by covalent bonding, and if so, which electrons would be shared between the atoms.

NOTE: After the class has specified the electrons involved, point out that there is no real significance in placing the unpaired electron of

chlorine in the third p orbital. This is simply a convention used in writing the diagrams, and since all three p orbitals are equivalent in energy, any one of them could contain the unpaired electron.

4. Show, by re-drawing the diagram on the board, that this bonding would indicate an overlapping of the s orbital of the hydrogen atom with the p orbital of the chlorine atom that had only one electron. This sharing of electrons results in a bonding orbital for the two electrons being shared.

5. Have the students describe the resultant electron arrangement for each atom in the compound. Be sure that they perceive that each atom can now be considered to have a noble gas electron configuration.

E. Suggest that there is a simpler way to demorstrate the bonding between hydrogen and chlorine.

1. Once again, have the students state which electrons are involved in the formation of chemical bonds (valence electrons).

2. Using chlorine, as the example, ask someone in the class to specify what else is in the chlorine atom other than the seven valence electrons (nucleus and subvalence electrons).

3. Once eliciting sufficient response to this question, introduce the term Kernel (write it on the board) to designate the nucleus and all the electrons that are not in the valence energy level. Tell them that the kernel will be represented by the symbol for the element--Cl, in this case.

4. Correlate the placement of the electron-dots areound the symbol with the orbital boxes from the box diagram and draw a dot structure for the chlorine atom--:Cl . The positioning of the dots can be explained in terms of electron repulsion.

BEST COPY AVAILABLE

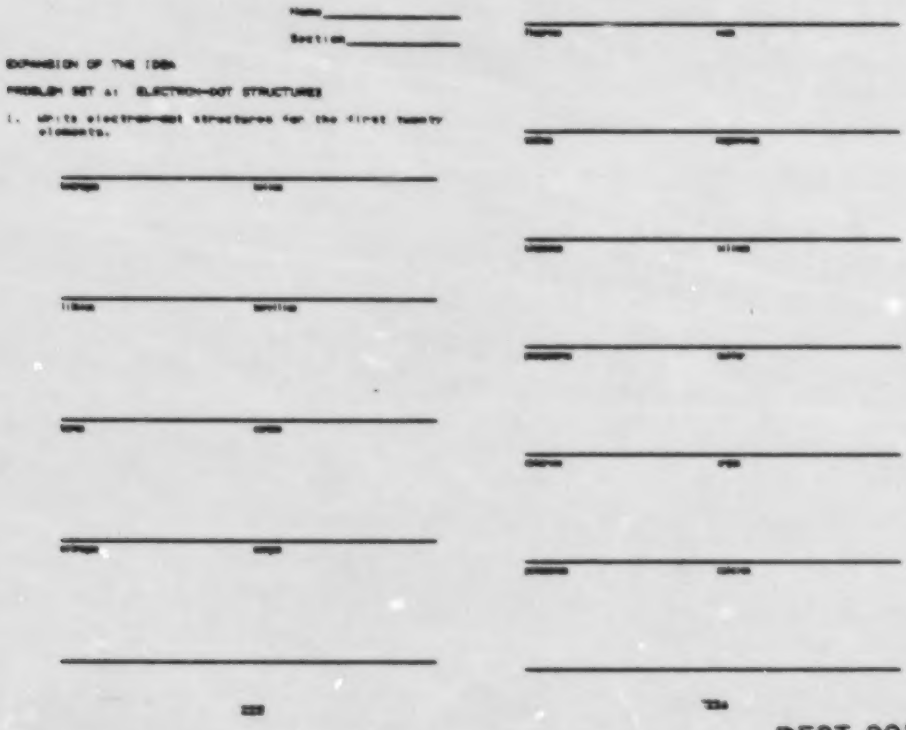
NOTE: The teacher should remind the students that these electrons will arrange themselves three-dimensionally around the atom, and this representation is not meant to imply that the pairs of electrons are ninety degrees apart.

5. Invent <u>electron-dot structure.</u> (Lewis structure)

6. Ask the class to review (orally) the significance of the parts of the dot structure.

7. Ask the class for input so that you can draw an electron-dot structure for hydrogen. Have them explain the parts of this structure. Show them how these two dot structures can be combined to yield an electron-dot formula for the hydrogen chloride molecule. Include a review of the noble gas electron configurations for the two elements involved.

F. Direct the students to the ELECTRON-DOT STRUCTURES worksheet, and go through the writing of electron-dot structures for the first ten elements. Student-teacher interaction should be maximized during this exercise.



BEST COPY AVAILABLE

1. When the dot structures for these first ten elements have been completed, present the class with the fact that III-A and IV-A elements often form three and four covalent bonds, respectively.

2. Ask the students to look at their box diagrams for boron and carbon and to try and decide how they could form three and four covalent bonds, respectively.

3. Hopefully, someone will observe that there are empty orbitals available in these elements (empty p-orbitals). If not, the teacher may finally need to present the idea that one of the valence s-electrons is promoted to an empty p-orbital--under the disturbing influence of other atoms forming bonds--and, at the same time, the energies of the p-electrons adjust so that the three or four single electrons now all have equivalent energies. When that happens, two additional bonding "sites" or locations are created at which covalent bonds can be formed. Invent <u>hybrid orbitals,</u> as s and p orbitals in a bonding atom that have the same energy.

4. Refer the students to the original dot structure that was drawn for carbon (when the initial discussions of methanol were held). At that time, we assumed the situation that we have now explained in terms of a more refined model of the atom. Tell the class that either of the dot structures for these families of elements will be acceptable, but the ones with all single electrons probably show the bonding situation more clearly. (The other dot arrangements can be called <u>ground state</u> electron arrangements.)

5. After the foregoing explanation for III-A and IV-A dot structures is presented, ask the class if this situation would also apply to V-A elements. Have them support their responses by using box diagram information. (No, there are no empty valence orbitals in nitrogen.)

G. ASSIGNMENT: When the electron-dot structures for the first ten elements have been drawn, assign the remainder of the exercise sheet as homework. The assignment includes the Problem Set a and Problem Set b. READING a: ARE ELECTRONS SALTY?, can be assigned at this time, along with the questions related to it.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

PROBING BY QUESTIONS

1. Using the model for metallic bonding developed in this reading, offer an explanation for the good heat conductivity characteristics of metals.

1. If the electrons are delocalized and belong to the metal as a whole, they are not restricted in their movement. Assuming the electrical conductivity in solids is directly related to the ease of electron movement, we can explain the high heat conductivity of metals from our model.

2. As temperature increases, the electrical conductivity of metals decreases. Offer an explanation for this phenomenon in terms of the metallic bonding model.

239

2. Heat conductivity involves the transport of thermal energy. In ionic and covalent solids this is accomplished by slight vibrations of the ions or atoms in the positions they occupy. In metals the thermal energy can be transported through the solid metal by the very mobile electrons. This mobility is hindered as the vibrational motion of the metal cations in the lattice increases. The vibrational motion of the metal ions increases with temperature. This explains why the electrical and heat conductivity of metals decrease with increasing temperatures.

EXPANSION OF THE IDEA

PROBLEM SET 5: ELECTRON-DOT STRUCTURES

3. Write electron-dot structures for the following molecules.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

 BEST COPY AVAILABLE

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

IV. COVALENT BONDING: FOURTH DISCUSSION

A. These discussion notes deal with the conceptual inventions of <u>single, double, and triple</u> covalent bonds and <u>electronegativity.</u>

B. This discussion should begin with a thorough review of the homework assignment given at the end of the third discussion.

1. Ask how the dot structures of the ten elements, done as homework, compare to the ten that were done in class (they are the same).

2. Ask if the dot structures for <u>A-column elements</u> in the remainder of the fourth period, the fifth period, and the sixth period should have dot structures similar to the ones that were drawn. Be sure that the students understand that they can generalize the dot structures for a given family of A-column elements to all the elements in that family.

C. Review the structures for the <u>compounds</u> given, as they are requested by the class.

1. The structure for carbon dioxide may have given them some trouble, and if so, work out the structure on the board by arranging the oxygen atoms' diagrams so that no distortion of carbon's dots occurs.

NOTE: This is only a temporary concession, as the appropriate modifications will be made after the multiple bond inventions.

2. Move on to any question that may have arisen in the writing of the dot structures for the diatomic elements. For example, there is no special position for the pair of dots in the dot structure for the nitrogen atoms. The relative positioning of pairs of dots and single dots, in the initial drawing of the structures for the elements, may have led to some confusion when the students tried to make the atoms bond together.

NOTE: Although the oxygen molecule does not contain a double bond, the structure showing a double bond is consistent with the work being done at this time (that is the mutual formation of noble gas electron configurations for the elements in the molecule). After the students understand the concepts being presented here, the teacher could reveal this fact to

the students, if it seems desirable to do so.

3. When questions on the drawing of the structures have been satisfactorily answered, center the attention of the students on the number of electrons being shared between any two atoms in the various molecules. As a result of this discussion, invent <u>single bond, double bond, and triple bond.</u>

4. Show how lines can be used to represent a pair of bonding electrons (for example, H-Cl:).

5. A good question to ask the students is, "Why are the seven diatomic elements diatomic?" (They reaction with each other to achieve noble gas electron configurations.)

D. A question should now be raised about the relative amount of attraction that two atoms will have for a pair of electrons involved in a covalent bond.

1. Have the students suggest factors that will affect how strongly an atom can attract the pair of electrons. (These factors should include such things as the nuclear charges on the atoms involved and the tendency of an atom to gain electrons.)

2. Ask the students which elements, in general, would have the greater tendency to gain electrons. They should be able to answer this from information they have gained from the discussions of ion formation of the A-column elements.

3. Refer the students to their first ionization energy graphs and have them relate ionization energy (high or low) with the tendency of an atom to gain electrons. Note that this relationship breaks down for the noble gases because they have completed s- and p- valence sublevels.

4. Invent the name <u>electron affinity</u> for the tendency of an atom to gain electrons. Have the students tell (generally) which atoms would have high electron affinities and which atoms would have low electron affinities.

5. Be sure that the students understand that the elements with the higher electron affinities (and higher ionization energies) will attract the electrons in a chemical bond more strongly than atoms with low electron affinities (and low ionization energies).

E. Tell the students that a scale has been developed to measure the relative attraction that an atom has for a pair of electrons in a bond.

1. Invent the term <u>electronegativity.</u>

2. Call the attention of the students to their handout sheet of a table of electronegativities.

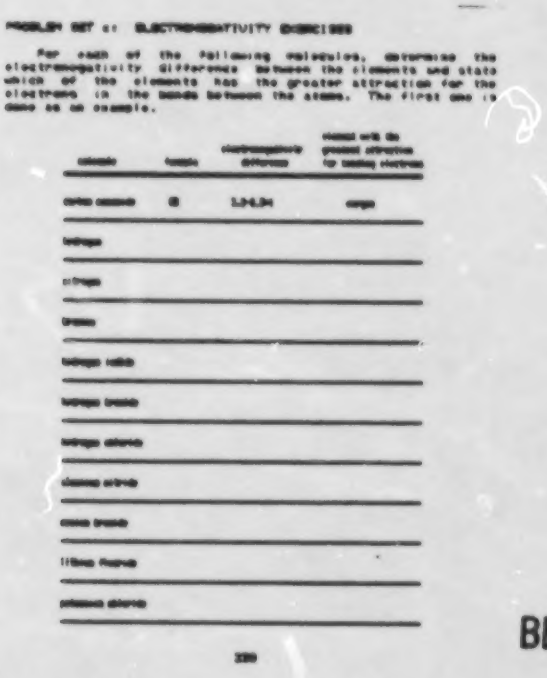
F. Ask the students to examine these electronegativity tables carefully. Allow them a few minutes to familiarize themselves with the table.

1. Ask which family of elements has the lowest values for electronegativity and which family of elemets has the highest values (the noble gases and the VII-A elements, respectively).

2. Discuss the trends in values for A-column elements and the absence of trends for the B-column elements.

3. Ask them to determine the range of the electronegativity scale. When they have stated a range, ask them why the noble gas family has been left off the table. The range of values that should finally be established is 0 to 4.

G. Hand out Problem Set c: ELECTRONEGATIVITY EXERCISES.



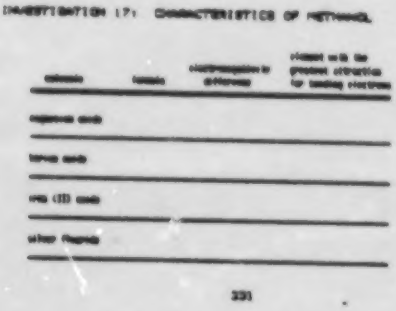
PROBLEM SET c: ELECTRONEGATIVITY EXERCISES

For each of the following molecules, determine the electronegativity difference between the elements and state which of the elements has the greater attraction for the electrons in the bonds between the atoms. The first one is done as an example.

	B	3.0-2.0 = 1.0	oxygen

339

These sheets should be completed by the next class meeting as they will be used in developing the concept of polarity.



INVESTIGATION 17: CHARACTERISTICS OF METHANOL

331

BEST COPY AVAILABLE

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

V. COVALENT BONDING: FIFTH DISCUSSION

A. This discussion deals with the relationships between the electronegativities of elements in a compound and the type of bonding negativities of elements in a compound and the type of bonding (ionic or covalent) between the elements. Included within this discussion will be the conceptual inventions of <u>polarity, non-polarity</u> and the form for representing ionic compound dot structures.

B. Have the class describe (by way of review) how an <u>ionic</u> bond is formed between two elements--starting with neutral atoms.

1. Suggest that they use the formation of potassium fluoride, as an example.

2. As the description is given the teacher should draw a dot structure representation of these elements and the resulting ionic compound on the board.

<u>example:</u>

3. Tell the class that this is a method for drawing a dot structure for ionic compounds.

C. Ask why it is reasonable to assume that this substance is ionic. Lead the discussion so that at least all of the following reasons are given. (It is desirable to write the reasons given on the board, as they are given by the class or developed through discussion.)

1. The substance forms ions in water solution (conducts electricity).

2. Both elements attain noble gas electron configurations, if ionic bonding is assumed.

3. The non-metal has a much greater attraction for electrons (higher electronegativity) than the metal.

D. Have the students compare the electronegativities of the two elements in the compound.

1. Write the two electronegativity values, and the difference between them, on the board.

E. Ask the students to choose a molecule (from the ELECTRONEGATIVITY EXERCISES sheet) that would be covalently bonded.

1. Lead them through the determination of the electronegativity difference for this molecule, and then center their attention on the comparative electronegativity differences for the ionic and covalent substances.

F. Ask what an electronegativity difference of zero would indicate about how the atoms in the molecule are attracting the pair of bonding electrons. (The idea to be developed here is simply that the atoms share the electrons equally.)

G. Ask about the relative attraction involved in a covalent molecule, where the difference in electronegativity is not zero.

1. Lead the students to the idea that, as the electronegativity difference increases, the compound becomes more ionic in character.

2. As an example of the simultaneous ionic and covalent character of a substance, refer to hydrogen chloride (HCl).

a. Show how it is reasonable to describe HCl as a covalent substance, in terms of electron arrangements of the elements (noble gas configuration).

b. Use the fact that HCl ionizes in water solution to show that it also has properties of an ionic substance.

H. Ask the class if it is possible that KF could have some covalent character in its bonding.

NOTE: In the discussion of this question, the teacher should insure that the class develops the idea that covalent character exists in all bonds, with some bonds being more ionic than covalent.

I. When the class has accepted the foregoing, the teacher should establish the convention for determining whether a binary substance is ionic or covalent, using electronegativity difference.

1. According to the way the electronegativity scale is established, an electronegativity difference of 1.7 corresponds to about 50% ionic character.

2. Using the convention, one refers to a compound as <u>ionic, if the electronegativity difference is greater than or equal to 1.7.</u> A compound is considered to be covalent, if the electronegativity difference is less than 1.7.

NOTE: It need not be a matter of discussion, at this time, that this rule only applies well to two-atom units. This point can be made later, when more complex substances are considered.

J. Ask the students to use this convention to label the substances, listed on the ELECTRONEGATIVITY EXERCISES sheet, as <u>ionic</u> or <u>covalent.</u>

K. Using HCl as an example, have the class review the relative attraction for the bonding pair of electrons in the molecule.

L. Lead the class to the conception of a partial negative charge for chlorine and a partial positive charge for hydrogen, due to the difference in the ability of the two atoms to attract the pair of electrons.

M. Invent <u>polarity.</u>

N. Lead a similar discussion for hydrogen (H_2) and invent <u>non-polarity.</u>

O. Ask the class if ionic substances would be considered polar or non-polar. Have them support their answers in terms of their knowledge. (They would be considered as being at the extrememe of polarity.)

P. Tell the students to label the substances, listed on the ELECTRONEGATIVITY EXERCISES sheet, as polar or non-polar.

NOTE: When this labeling is completed, the teacher can call the attention of the students to the labels as <u>non-polar covalent, polar covalent,</u> and <u>polar ionic.</u> Point out that the expression <u>polar ionic</u> is redundant because ionic substances are necessarily polar.

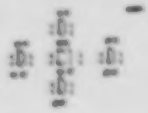
Q. This would be a good time in the discussions to review the concepts developed to this point.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

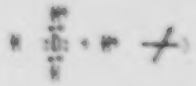
R. Assign Reading 5: ARE ELECTRONS COORDINATED? and its questions.

[illegible]

1. Start with Cl^- ion and each oxygen atom forms a coordinate covalent bond.



2.



BEST COPY AVAILABLE

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

VI. COVALENT BONDING: SIXTH DISCUSSION

A. Draw a line diagram for hydrogen chloride on the board (H-Cl:). Ask the class to orally review their learning about this substance (for example, it will ionize in water solution, covalently bonded, polar, etc.).

1. Have them specify which part of the molecule is partially negative and which part is partially positive.

2. Draw an arrow to represent this.
(example: --/--))

NOTE: This is a common way to represent the location of charges in dipoles.

3. Invent the term dipole to apply to a particle that has distinct positively and negatively charged regions. (Write it on the board.)

4. Point out that polar molecules, such as HCl, contain permanent dipoles. Have the class review the reason why such dipoles would be permanent.

B. Ask the class to instruct you as you draw a line diagram for carbon dioxide on the board $\ddot{O}=C=\ddot{O}$.

1. Center the attention of the class on the bonds between one of the oxygens and the carbon. Ask if the bonds between the carbon and oxygen would be polar or non-polar. Have them give supportive evidence for their response. (polar).

2. Ask how the dipole arrow should be drawn and then draw it.

3. Repeat the procedure for the bonds between the carbon and the other oxygen. The drawing on the board should look like:

4. Ask what shape the carbon dioxide molecule would appear to have. Invent the term linear molecule, after they have responded.

5. Tell the students that it has been determined, through experimentation, that the carbon dioxide molecule is <u>non-polar.</u> Ask them to propose an explanation for this experimental observation.

NOTE: The teacher should decide if a brief discussion of vectors is appropriate at this time. The important concept for the students to have is that dipoles of equal magnitude and opposite direction (within a molecule) will cancel the polar effects of the individual bonds.

6. Ask the class to predict the shapes of some diatomic covalent molecules, such as H_2, HCl, AlN, etc. Be sure that they understand that the shape of the molecule can be ascertained by imagining a line drawn through the nuclei of the atoms in the molecule.

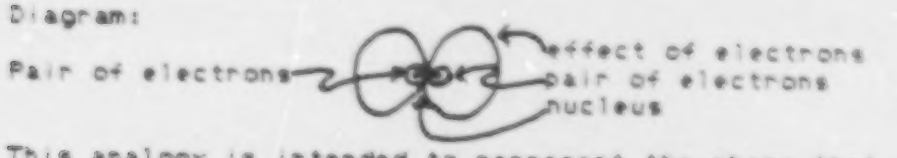
C. Put an acetate on the overhead projector. Have four flat, circular magnets available to use for this demonstration.

D. Put one iron bolt with a hexagonal head on the overhead projector to represent the nucleus of an atom. Ask the class how <u>two</u> magnets (placed on the overhead so that they are attracted to the bolt) would interact with each other. (repel) Perform this operation. (It will probably be necessary to vibrate the bolt to get the magnets to a linear orientation.)

NOTE: It is important that the students realize that this is merely an analogy. We would not expect one pair of electrons to attract another pair of electrons, if they were properly aligned.

1. Using an overhead marker, draw around the "electron pair" magnets to indicate the effective space of the electron pairs and their orientation with reference to each other.

Diagram:



This analogy is intended to represent the shape that two pairs of valence electrons would assume.

2. Ask the class to predict the arrangement of three pairs of valence electrons around a nucleus. Add another magnet and then draw around the magnets with the overhead marker. (Use a different area of the acetate.)

3. Ask the class to state the angles between the pairs of electrons in each of the foregoing cases.

4. Have the class predict the arrangement of four pairs of electrons around a nucleus. After their predictions have been made, hold the "nucleus" with one finger and bring the other magnet up to it.

NOTE: This procedure will be unsuccessful, and one or more magnets will jump out of place. You may want to repeat your attempts two or three times for emphasis. One option for doing this procedure is to invert a small beaker over the system of three pairs and a nucleus and then bring in the other "pair." As this is done, one of the magnets will be ejected from the system.

5. Ask the class if it is possible to have four pairs of valence electrons around a nucleus. When they have agreed that there can be four pairs and have given examples to support this fact, ask what this demonstration tells them about the arrangement of four pairs. (The four pairs are not all in the same plane.)

E. Referring to the drawings of the effective spaces occupied by the pairs of electrons (on the overhead), suggest to the class that inflated round balloons might suggest a more accurate model for the arrangement of electron pairs around a nucleus.

NOTE: You will need five balloons for this demonstration--two pairs of balloons, made by taping the mouth pieces together, and one single balloon.

1. Use a pair of balloons to illustrate the shape assumed by two pairs of electrons. Point out that the shape is linear (a straight line or a plane could be imagined as going through the center of the nucleus and the two pairs of electrons).

2. Combine the single balloon with the pair of balloons and get the class to agree that the planar arrangement indicates the least amount of repulsion among the three "pairs" of electrons.

3. Interlock the two pairs of balloons and then flatten them out on the demonstration table. Let them go to allow them to reach a "most comfortable" state. Lead the class in a discussion of this shape. Develop the idea that only two "pairs" and the "nucleus" could be considered to fall into a plane. The other two pairs are outside this imaginary plane.

4. Invent the term <u>tetrahedral</u> for the shape assumed by the four pairs of electrons.

NOTE: The teacher will probably need to spend some time going over this shape with the students. One aid that is easy to prepare consists of four identical equilateral triangles (cut from pasteboard). These can be made of sufficient size to enclose a ball-and-stick model corresponding to a nucleus and four pairs of electrons.

5. Define the angle (109.5°) between the pairs of electrons.

F. The teacher can now establish a method for representing the tetrahedral structure on paper.

1. Review the fact that only two of the pairs and the nucleus will be in a plane. You may wish to use a ball-and-stick model to show this.

2. A representation of this idea may be drawn on the board, as follows:

3. One of the remaining two pairs can be represented as coming <u>out</u> of the board (◢) and the fourth pair can be represented as going <u>into</u> the board (⋘). The total drawing for the tetrahedral arrangement of four pairs of electrons would then be:

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

G. Assignment: Assign Problem Set d: ELECTRON AND ATOM ARRANGEMENTS as homework.

1. Describe the shape represented by the drawings in Question 1.

2. Tetrahedral, with an angle of 109.5° between any two pairs of electrons and the nucleus.

VII. COVALENT BONDING: SEVENTH DISCUSSION

A. In this discussion the idea of polarity and non-polarity of molecules will be extended by using examples from the homework assignment, ELECTRON AND ATOM ARRANGEMENTS. Demonstrations that show how polar or non-polar substances interact with a static charge will also be used.

B. Refer the class to Question 3b, on their homework (carbon tetrachloride). Ask for a volunteer to draw the diagram on the board.

1. Ask if the individual C-Cl bonds are polar or non-polar. Require supporting information for the answer. (Polar--difference in electronegativity.)

2. Ask if the CCl_4 molecule is polar or non-polar. Regardless of the response to this question, the teacher should lead a discussion to rationalize the non-polarity of the substance. A ball and stick model representing the compound may be useful in explaining the dipole directions and the cancellation of dipolar effects.

3. Instruct the class to label their diagrams for CCl_4 as <u>non-polar,</u> after they have concurred with the explanation given in the discussion.

C. Center the attention of the class on their diagrams for ammonia (Question 3c).

1. Ask for a student volunteer to draw the ammonia diagram on the board. Ask that student (or another) to use the drawing to explain the polar or non-polar character of the substance (at the board). (Ammonia is a polar molecule.)

BEST COPY AVAILABLE

2. When the explanation is completed, tell the students that the angles formed between the nitrogen and hydrogen nuclei in ammonia are determined by experiment to be 107° and not 109.5°, as they are in carbon tetrachloride. Ask them to study their diagrams and to offer an explanation for this deviation from the expected tetrahedral angle.

NOTE: If the students are totally stumped by this question, the teacher can get them started by suggesting that they make a list of the differences between molecules of ammonia and carbon tetrachloride. Allow some time for the students to think about this question. This should lead to responses from the class that will be valuable to the development of the idea of unequal repulsion between bonding and non-bonding pairs of electrons. Pursue the discussion until the reasoning is clear to the students.

3. Review the shapes of the carbon tetrachloride and ammonia molecules: tetrahedral and trigonal (three-sided) pyramidal, respectively.

4. Instruct the class to label their diagrams for ammonia as <u>polar.</u>

D. Ask the class to state the shape of the water molecule (bent). Draw a diagram for water on the board, using the information given by the class. Have the students specify whether water is polar or non-polar.

1. Ask how they would expect the angle between the hydrogen atoms, in water, to compare with the bond angles in CCl_4 and NH_3.

2. At the point in the discussion that seems appropriate, establish that the bond angle is 104.5° for water.

E. The materials needed for the following demonstrations using static charges are as follows:

1. Three polyethylene wash bottles: one containing distilled water, one containing methanol, and one containing carbon tetrachloride.

2. Three 400ml beakers to catch the liquids.

3. A glass rod and a rubber rod.

4. Silk and wool (or dry hair).

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

F. Review with the class the overall charge characteristics of water (that is, neutral, polar, contains permanent dipole).

G. Rub one of the rods to create a static charge and then have a student volunteer squeeze a steady stream of water from the wash bottle into one of the beakers. The bottle should be held about 30-40cm above the beaker.

H. Bring the charged rod near the stream of water. Ask the class to explain the observed phenomenon. A diagram on the board of the alignment of water molecules would be useful to reinforce the explanation given by a student.

I. Ask if similar results would be expected, if an oppositely-charged rod were used. Perform this demonstration.

J. Repeat the questioning and demonstration using carbon tetrachloride.

K. Have the class make predictions about the behavior of methanol near the charged rods. After they have given reasons for their predictions, repeat the demonstration using methanol.

L. ASSIGNMENT: Have the class make a diagram of methanol, including the dipole "arrows," to show that methanol is a polar molecule.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

INTRODUCTION

To develop general solubility rules based on structural consideration, this investigation builds on the ideas of molecular shape and polarity attained during the COVALENT BONDING discussions. This leads to the conceptual invention of van der Waals forces which are inter-molecular forces of attraction. The solubility concept will be derived directly from laboratory data, but the intermolecular force concepts will be developed using written work, class discussion, and readings as starting points.

CONCEPT

Characteristic physical properties of substances are related to the molecular or ionic structure of those substances.

EQUIPMENT

A. Labware

13x100 mm test tubes
corks to fit test tubes
test tube racks
plastic spoons

B. Reagents

iodine crystals
potassium chlor de
hexane
carbon tetrachloride
distilled water

PRE-LAB INSTRUCTIONS

A. The liquids can be conveniently supplied in and dispensed from plastic wash bottles. Alternatively, the liquids can be supplied in large erlenmeyer flasks, and 10 ml pipets with bulbs can be used for dispensing the liquids.

B. The iodine and potassium chloride can be set out in covered, labeled petri dishes. The iodine crystals should be broken into small pieces to encourage the students to use small amounts.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Name_____

Section_____

INTRODUCTION

As the title of this investigation suggests, you will be studying the relationship between the structure of various substances and some of their physical properties--specifically, solubility, melting point, and boiling point. This study will allow you an opportunity to utilize your learning about covalent bonding.

EXPLORATION

1. Use approximately 2 ml of the liquid and a small amount of solid. When combining two liquids, use equal volumes of the two liquids. Stopper and shake the test tubes vigorously.

CAUTION: If you get any of the solids or liquids (other than water) on your hands, wash your hands with soap and water.

2. Label the various test tubes for the tests and keep them and their contents to compare with the other tests done. Such comparisons will help you to decide about the solubility of various substances.

3. Describe the results of your solubility tests in the appropriate spaces.

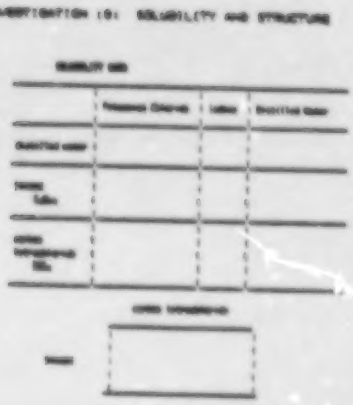
4. When you have completed all the tests, dispose of the contents of the test tubes according to the instructions of your teacher.

1. Point out the location of the solids and liquids in the laboratory. You may wish to include comments on care in handling the chemicals.

   Stress the use of <u>small</u> amounts of solids. Point out that the volumes listed are approximate.

4. A large brown bottle can be placed under the ventilation hood and used as a waste liquid container. This method of disposal would be preferrable to pouring these liquids down the drain.

BEST COPY AVAILABLE



CONCEPTUAL INVENTION (THE IDEA)

Acceptable Answers

1. POLAR

   potassium chloride

   water

   NON-POLAR

   iodine
   hexane
   carbon tetrachloride

2a. Hexane can be classified as non-polar from two different lines of reasoning. The structure can be drawn out (or a model built) and the dipole directions analyzed. Alternatively, the data from the solubility tests can be compared, and the similarities of behavior with other known non-polar substances can be noted to lead to the conclusion that hexane is non-polar.

3. Polar solutes dissolve in polar solvents; non-polar solutes dissolve in non-polar substances.

BEST COPY AVAILABLE

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

INVENTION DISCUSSION

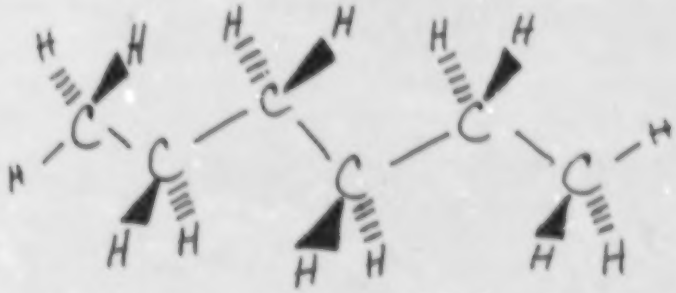
A. Review the listings (from Question 1) for the substances used in the investigation (that is, whether they were polar or non-polar).

B. If questions arise from students about the polarity of hexane, these questions can be used to begin the discussion. If not, the following guidelines are suggested for discussing hexane.

NOTE: It would be useful to have model built (ball and stick, plastic, etc.) of an n-hexane molecule.

1. The n-hexane model can be placed on the stage of an overhead projector so that all the carbon atoms are in one horizontal plane. The focus of the projector can then be varied to show that the atoms in the molecule fall into one of three horizontal planes--one on the surface of the projector and the two above the projector surface.

2. When this structure is understood in terms of the location of the atoms relative to each other, a diagram of the molecule can be drawn on the board as follows.



Have students suggest the directions of dipoles, draw them in on the diagram and lead the class to an understanding of the non-polar character of hexane.

3. Proceed in a similar manner, using the water molecule, to re-establish that water is a polar compound.

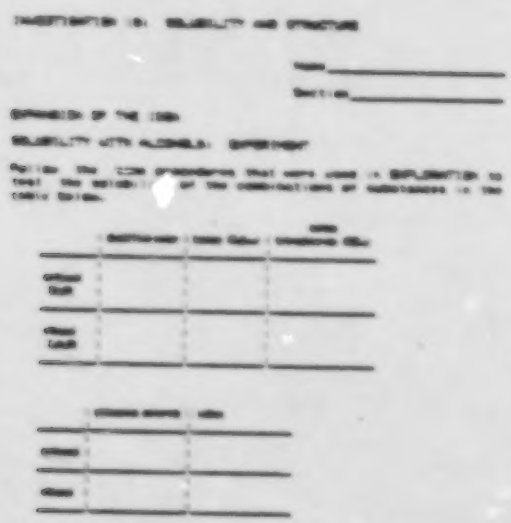
D. Ask students to read the general rules that they formulated to answer Question 3. There is ample information from the data to allow the general statement desired here.

If students have been unable to write such a statement, the teacher will need to lead a discussion of the data to help the students arrive at an adequate statement.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

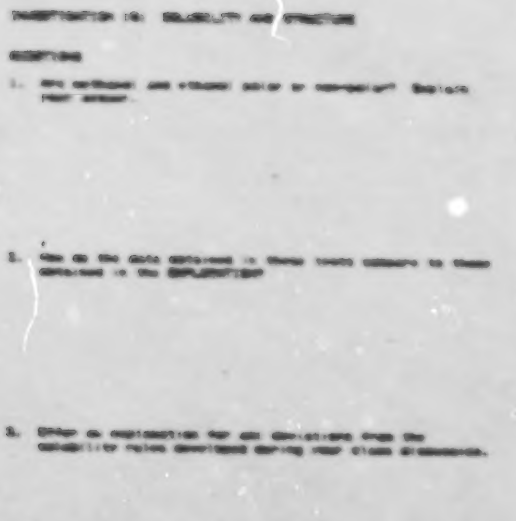
EXPANSION OF THE IDEA

A. Experiment



1. All combinations in this part of the table are soluble to some degree.

   Iodine dissolves in methanol and ethanol, but KCl is insoluble in these solvents.



1. Polar Dipole vectors drawn on the molecules would result in a net dipole for the molecules.

2. There are some compounds that seem to contradict the other data. The alcohols (polar) dissolve polar and non-polar substances. KCl (polar) is insoluble in the alcohols.

3. The alcohols have regions of their molecules that resemble non-polar substances. These parts of the molecules could give rise to compatible attractions between the alcohols and non-polar solutes.

BEST COPY AVAILABLE

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Discussion of Solubility with Alcohols Experiment

1. Begin the discussion by working through the questions which follow the activity:

   a. Ethanol can be analyzed and be determined to be polar in a way similar to the determination for hexane. The obvious difference, of course, is that ethanol is polar.

2. Review the substances which provided exceptions to the solubility rule. Ask the students to offer explanations for the large solubility of iodine (non-Polar) in methanol (polar).

   a. It might be helpful to have built a model of the methanol molecule to show the class.

   b. Ask the students how I_2, for example, dissolves in methanol. Allow time for the students to think about the question and formulate responses.

   c. A model for this solubility could involve the interaction of the methyl group and (note its similarity to that of the methane or carbon tetrachloride molecules) with the non-polar, iodine. There are, of course, other models, such as the formation of the soluble I_3^- ion by some mechanism. This second model seems especially feasible in light of the color differences between the iodine solutions with the non-polar solvents and the iodine solutions with the alcohols.

B. When the students have the idea that structure is related to these other physical properties, assign Reading a, BETWEEN THE MOLECULES...AN ATTRACTION STORY.

C. Assign Problem Set 18, MELTING POINT, BOILING POINT, AND STRUCTURE at this time.

[illegible]

Name_____

Section_____

[illegible]

PROBLEM SET: MELTING POINTS, BOILING POINTS, AND STRUCTURE

[illegible]

[illegible]

1. For each of the groups of compounds in the table, list [illegible], differences, and any trends that you can observe among the characteristics listed.

[illegible]

Students should be able to complete the table by writing the remaining formulas, the molecular weights, and the polarity characteristics of these substances.

1 and 2.

There are several similarities and difference that can be listed. The discussion of these is challenging and interesting. One of the most significant trends (with exceptions for H_2O and HF) is that m.p. and b.p. increase as molecular weight increases. (This is an acceptable response to Question 2.)

BEST COPY AVAILABLE

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

3. Increased weights mean greater numbers of electrons, and, in many cases, physically larger molecules. These factors give rise to stronger van der Waals attractions (see Reading a).

4. These exceptions are due to the very high polarity of these molecules, which gives rise to strong electrostatic attractions between molecules (see Reading b).

5. All are ionic, but their differences in electronegativity decrease in order from NaF to NaI.

6. Their m.p.'s and b.p.'s would be much higher. The attractions between molecules (or ion pairs) in these substances are essentially ionic attractions, whereas the intermolecular attractions of the other substances are van der Waals types of attractions.

7. As the ion pairs become less polar, their attractions for other ion pairs decreases. These pairs more closely approximate covalently bonded molecules, as the polarity decreases.

BEST COPY AVAILABLE

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

DISCUSSION OF PROBLEM SET 18

The questions on the problem sheet should be discussed and areas of confusion cleared up.

1. Question 4 gives a good opportunity to lead to the idea of hydrogen bonding.

2. Ask the class what each of these compounds has in common. (hydrogen)

3. Ask the students to explain whether the presence of hydrogen in the molecules would account for the properties. The mere presence of hydrogen could not account for the properties because all of the compounds in Table IV contain hydrogen.

4. Since there are only two elements in each compound, the unusual properties must be related to the other elements. Ask if nitrogen, oxygen, and fluorine would be able to attract electrons well. (Yes, they have high electronegativities.)

5. Ask how the electronegativities of these elements compare to that of hydrogen (higher).

6. Ask why a strongly polar, covalent molecule with hydrogen as one of the elements would provide for more intermolecular attraction than a molecule that was equally polar but did not contain hydrogen. (Hydrogen has only one electron and, as it is pulled away by the attraction of a more highly electronegative atom, the nucleus of the hydrogen atom becomes exposed. This direct exposure of the nucleus allows for a stronger attraction for another molecule's negative charge.

7. Invent <u>hydrogen bonding</u> to name the phenomenon just described. Repeat the fact that this is an <u>intermolecular</u> attraction that occurs in molecules containing hydrogen and small, highly electronegative elements such as nitrogen, oxygen, and fluorine.

8. Assign Reading b, HYDROGEN BONDING, at this time.

9. Ask the students to apply what they have learned to compounds at the extreme end of the polarity spectrum--ionic compounds. In these sections we will be making an assumption that we can properly discuss these compounds as <u>molecules</u>. That assumption does not seem to be too gross in light of ion pair formation in vapor states. The facts of high melting points and boiling points or these compounds remain, and the main reason the values for those properties are so large is that there is very strong attraction between (and among) the ions.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

a. The values for the melting points and boiling points of these compounds are as follows.

compound	m.p. (°C)	b.p. (°C)
NaF	988	1695
NaCl	801	1413
NaBr	755	1390
NaI	651	1304

b. A reasonable explanation for the decrease in m.p. and b.p., when the stubstances are listed as they are above, is that the substances tend to be more covalent as one moves down the list (that is, there is less and less difference in electronegativity).

INVESTIGATION 19: PERIODICITY AND REACTIVITY

TEACHER'S GUIDE

INTRODUCTION

The content of this investigation and its placement in the course allow a broad range of idea expansions to be made during discussions of the activities. Trends in reactivity and possible reaction mechanisms of I-A and II-A metals as well as VII-A non-metals are investigated. Also, the general reactivity trends of A-column metals and non-metals will be studied. The I-A metals experiments are performed by demonstration for reasons of student safety. Following the I-A metals demonstrations, this investigation continues with student-performed experiments dealing with the reactivities of the II-A and III-A metals and the VII-A non-metals.

CONCEPT

The chemical properties (reactivity) of metals are determined by their electron structures.

EQUIPMENT

A. Labware

1. For I-A metals

a. 400 ml beakers each containing 50 ml of dry kerosene
b. glass petri dishes or watch glasses to cover 400 ml beakers
c. forceps
d. knife
e. deflagrating spoon with wire screen to cover the bowl
f. 400 ml beakers each containing 50 to 100 ml of distilled water
g. 18x150 mm test tube with rubber stopper
h. paper towels
i. matches

2. For II-A and III-A metals

a. 50, 100, 150 ml beakers
b. laboratory burners
c. ringstands with rings
d. wire gauze with asbestos center
e. 13x100 mm test tubes
f. corks to fit test tubes
g. test tube racks

B. Reagents

1. For I-A metals

a. pea-sized pieces of lithium metal-4 (in beaker with kerosene)
b. pea-sized pieces of sodium metal-4 (in beaker with kerosene)
c. pea-sized pieces of potassium metal-4 (in a beaker with kerosene)
d. pink and blue litmus paper

2. For II-A and III-A metals

a. large aluminum shot
b. 25 cm strips of magnesium ribbon
c. calcium turnings
d. HCl solution (500 ml of concentrated HCl added to 1500 ml of distilled water)
e. distilled water
f. pink and blue litmus paper
g. pea-sized pieces of lithium metal

NOTE: To avoid possible contamination of large quantities of reagents, it would be advisable to put the calcium, magnesium, and aluminum metals in smaller containers (for example, large evaporating dishes or smaller reagent bottles) in a centrally located place in the laboratory.

3. For VII-A elements

a. 0.1M solutions of NaF, NaCl, NaBr, and NaI-100 ml each
b. Hexane in brown dropper bottles-100 ml
c. 100 ml of chlorine water in brown dropper bottles. Chlorine water can be obtained in several ways. Some of them are:

1) chlorine water purchased from a chemical supply house
2) concentrated HCl added to a commercial chlorine bleach until the gas formation ceases. This operation should be performed under a ventilation hood
3) A chemical reaction that produces Cl_2 gas. This gas can then be bubbled through distilled water. This operation should also be performed under a ventilation hood.

d. 100 ml of bromine water in brown dropper bottles. Some ways of obtaining the bromine water are:

1) Bromine water purchased fom a chemical supply house.
2) Pour elemental bromine into a flask containing distilled water, stir, allow the excess bromine to settle and decant the bromine water. The excess Br_2 can be used to pepare more bromine water.
3) dissolve 40 g of KBr in 500 ml chlorine water and call it bromine water.

e. 100 ml of iodine water in brown dropper bottles.
f. This solution can be prepared by dissolving 10 g of KI in distilled water, adding 5 g of solid iodine and diluting the solution to 500 ml.

NOTE: The halogen-water solutions need to be kept tightly covered, when not in use. If these solutions are purchased, it would be wise to order pint-sized containers to avoid wasting large amounts of the solutions. A convenient method for dispensing these solutions is to put them in brown 250 or 500 ml dropper bottles equipped with droppers that screw on the bottle. Depending on the availablity of dropper bottles and the size of the classes, it might be good to have two sets of dropper bottles containing the halogen water solutions and the hexane.

PRE-LAB INSTRUCTIONS

A. This set of demonstrations will be more effective if they are performed at a central table in the laboratory. This arrangement allows the students to observe the various interactions more closely and it also allows for freedom of movement when the students return to write their observations, hypotheses, and conclusions on their worksheets. The teacher should set an example for the students by wearing safety glasses during the demonstrations.

B. Pass out the first student sheets so that the students may fill it out at the appropriate time during the demonstration.

C. After explaining the materials to be used in Step 1 (that is, the beaker containing distilled water and sodium metal in kerosene), remove a piece of sodium from its container with the forceps and place it on a paper towel. Shave off some of the coating with the knife and have the students observe the metallic luster of the sodium.

1. Wipe off as much kerosene as possible with the paper towel and, using the forceps, place the sodium in the beaker of water. Cover the beaker immediately.

2. The students may request that you repeat the procedure. If you do so, it is wise to pour out the solution from the first reaction and get more water.

D. The procedure indicated in Step 3 is the collection and testing of the gas produced when sodium reacts with water.

1. Remove another piece of sodium from its container, dry it with a paper towel, and place it in the bowl of a deflagrating spoon. Secure a wire screen over the bowl of the spoon.

2. Fill an 18x150 mm test tube with water and invert it in one of the beakers of water.

3. Hold the sodium under the water and collect a test tube full of the gas. Do not test the gas for flammabilility until the students have had time to write their observations of the demonstration. Leave the spoon submerged until all the sodium has reacted.

4. Ask a student to help you test the gas for flammabilility. Allow time for observation writing.

5. Test the solution in the beaker first with blue and then with red litmus paper. Allow time for the results of the litmus tests to be written.

E. Instruct the students to complete Steps 3 and 4.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

INVESTIGATION 19: PERIODICITY AND REACTIVITY

Name_____

Section_____

INTRODUCTION

Steps 1 through 8 of this investigation refer to a set of demonstrations that your teacher will perform. In order to get the maximum benefit from those demonstrations, you will need to observe the systems very carefully. The main goal of this investigation is to ascertain information on the reactivity of the representative (A-column) elements.

EXPLORATION

1. Your teacher will place sodium metal in water. Observe carefully and record your observations.

2. Was the interaction a physical change or a chemical change? List evidence which supports your choice.

3. Your teacher will perform another procedure using sodium and water. Record your observations.

4. Write a balanced equation for the reaction between sodium and water.

1. After the demonstration, ask the students to follow the instructions in Step 1. Point out that Step 1 is the place for listing observations and Step 2 is the place for stating their hypotheses.

2. It is important that the students feel free to write a hypothesis in Step 2 along with what they consider to be supporting evidence. The teacher should allay any fears about writing hypotheses which may prove to be incorrect. Changes dictated by further activities should be added at the time that they are brought to light.

3. Instruct the students to wait to fill in Steps 3 and 4 until after the procedure using sodium and water is completed.

4. If some students have difficulty getting started with Step 4, suggest that they review the things that they have written about the interaction of sodium with water.

BEST COPY AVAILABLE

INVESTIGATION 19: PERIODICITY AND REACTIVITY

The students should be encouraged to hypothesize on the question in Step 5. After the students have finished writing on Step 5, perform the demonstration with lithium and water. This should be done in much the same way as the first demonstration with sodium was done, with the following additions.

1. The litmus paper tests should be performed on the solution that remains after the reaction.

NOTE: It may be advisable, depending on the size of the piece of lithium reacted and the amount of water in the beaker, to react a second piece of lithium in the same water prior to the tests with litmus paper. Trials of the demonstration before class will be helpful in determining the necessity for reacting a second piece in order to obtain a test easily observed with the litmus paper.

2. The students may want you to collect and test the gas that is formed. This collection can be done conveniently by holding an empty, inverted test tube above the piece of lithium while it is reacting.

5. How would the reaction rate of lithium in water compare to the reaction rate of sodium in water? Explain your predictions in terms of ionization energy and electronegativity.

6. Carefully describe the behavior of lithium in water, including a balanced equation for the reaction.

7. Predict the behavior of potassium metal in water, based on the information gained from the interactions of lithium and sodium with water.

8. Observations of potassium in water. (Include a balanced equation for the reaction.)

6. Allow time for the completion of the writing associated with Steps 6 and 7. It is important to point out to the students that their hypotheses about potassium in water should be based on the information obtained from the other reactions. Tell the students to write a balanced equation for the reaction.

8. Perform the demonstration with potassium in water. Since this reaction is the most spectacular of the three, you may be pressed to perform it a second time. If so, fresh water should be used. Test the resulting solution with litmus paper.

BEST COPY AVAILABLE

INVESTIGATION 19: PERIODICITY AND REACTIVITY

Prior to the students performing the laboratory work:

1. Review the location of the reagents in the laboratory and the precautions to be used when handling acid solutions.

2. Tell the students that more easily-interpreted observational data can be obtained if the physical configuration of the various metals is similar. For example, comparative rates of reaction for Ca and Mg can be better determined by using pieces of approximately the same shape and size. This can be accomplished by rolling magnesium ribbon into a tight coil before reacting.

3. Point out that unreacted metal pieces should be discarded in the trash cans and not in the laboratory sinks.

4. Alert the students to the need for writing correct formulas for compounds in order to get properly balanced equations for the reactions that occur. It is good to check the students' papers, as they are performing the experiments, to be sure that they have written the formulas properly.

9. Place one piece of aluminum, calcium and magnesium in separate beakers containing 30 to 40 ml of distilled water; allow them to sit for two or three minutes, and record your thorough observations for each. For best comparison try to get the metal pieces about the same size and shape.

10. Heating will sometimes speed up a reaction. If no change (or little change) is noticed after two or three minutes, heat the system and record your observations.

11. Write balanced equations for the chemical reactions observed.

12. Place out the beakers used earlier and put 30-40 ml of 6M HCl solution in them. Place the metals in the solutions as before and record your observations.

13. Write a balanced equation for each reaction of the metals used with the acid solution.

11. $Ca+2H_2O \rightarrow Ca(OH)_2+H_2$

13. $Mg+2HCl \rightarrow MgCl_2+H_2$

$2Al+6HCl \rightarrow 2AlCl_3+3H_3$

$Ca+2HCl \rightarrow CaCl_2+H_2$

NOTE: Students seldom list the possibility of calcium reacting with acid and water, but this idea can be proposed by the teacher later during class discussions.

BEST COPY AVAILABLE

INVESTIGATION 19: PERIODICITY AND REACTIVITY

CONCEPTUAL INVENTION (THE IDEA)

1. Rank the three I-A metals in order from least reactive to most reactive with water. Offer an explanation for the trend in reactivity (laboratory observations will not be sufficient to completely answer this).

   From least to most reactive the metals are lithium, sodium, and potassium, respectively. When one looks at the positions of these metals on the periodic table, the deduction can be made that their reactivity trend coincides with decreasing trends in ionization energy and electronegativity. These trends indicate that it should be easier to remove an electron from sodium than from lithium and from potassium than from sodium. Since these elements react by losing electrons, the easier the loss, the greater the reactivity.

2. How would the behavior of cesium in water compare to that of the other I-A metals used in this activity?

   Cesium metal should react quite vigorously with water. In fact, small amounts of cesium would probably react rapidly enough to cause an explosion. It would, therefore, appear to be more reactive than the three used in this experiment, but it would be closest in reactivity to potassium.

3. Provide reasons for this proposed behavior of cesium.

   The reasons for cesium's behavior are similar to the explanation given in Question 1. The ionization energy and electronegativity of cesium are lower than those of lithium, sodium, and potassium, as would be expected by those trends down a family of elements.

4. In general, do you think metals are more reactive with water or acid? Why should they be more reactive with one that with the other?

   Metals are generally more reactive with acids than with water. Possible reasons for this can vary quite a bit, but the most reasonable response will relate to the mechanism utilizing the hydrogen ion from the discussion of the I-A metals. Since the concentration of hydrogen ion is so much greater in acid solutions than in water, the rate at which the reaction would take place would also be greatly increased.

BEST COPY AVAILABLE

5. How would you expect Li, Na, and K to react with an acid solution? Explain your answer.

   The three I-A metals mentioned should react much more vigorously in acid solution than in water. The explanation, again, would center around the much greater hydrogen ion concentration in the acid solution.

6. Write the total electron configurations (not electron-dot diagrams for <u>neutral</u> atoms of Na, Mg, and Al.)

   Na $1s^2$ $2s^2$ $2p^6$ $3s^1$

   Mg $1s^2$ $2s^2$ $2p^6$ $3s^2$

   Al $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^1$

7. Write the total electron configurations of Na, Mg, and Al <u>after</u> they have reacted, including appropriate charges.

   Na^+ $1s^2$ $2s^2$ $2p^6$

   Mg^{2+} $1s^2$ $2s^2$ $2p^6$

   Al^{3+} $1s^2$ $2s^2$ $2p^6$

8. What had to happen to Na, Mg, and Al to change the configurations between Questions 6 and Question 7?

   The atoms had to have one, two, and three electrons (respectively) pulled off.

9. List as many relationships as possible between the position on the periodic chart and the reactivity for these metals (I-A, II-A, III-A).

   NOTE: It is not necessary to include all six metals in any given comparison, but all six metals should be included in your discussion.

   a. I-A metals are more reactive than II-A; II-A metals are more reactive than III-A.

   b. In the I-A family the metals are more reactive, as one moves down the family. (The same generalization could be applied to the II-A and III-A families and may be listed as separate relationships.)

   c. Reactivity of metals (A-column) decreases in going from left to right across a period (for example, Na, Mg, Al).

INVESTIGATION 19: PERIODICITY AND REACTIVITY

INVENTION DISCUSSION

A. Begin the discussion by reviewing the data obtained in Steps 1 through 8. Have the students dictate equations for the three reactions, as you write them on the board. Have them comment on the similarities among the equations and have them supply reasons for these similarities (for example, all equations balanced with the same coefficients, H_2 was formed in each case, base solutions resulted, etc.).

Equations for the respective reactions are as follows:

$2Na+2H_2O \rightarrow 2NaOH+H_2$

$2Li+2H_2O \rightarrow 2LiOH+H_2$

$2K+2H_2O \rightarrow 2KOH+H_2$

NOTE: A second reaction possibility for the potassium system is the burning of potassium in air to produce K_2O. Most students will not suggest this possibility, but it can be brought out by the teacher in the class discussion that follows.

1. Referring the students to Question 1 of THE IDEA, ask them to offer reasons, other than direct observation, to support the trend in reactivity. The responses written on the worksheets will dictate the best method for approaching this discussion. The emphasis here should be on the relative ease with which the metals forfeit their valence electrons to form positive ions.

2. A review of Questions 2 and 3 follows logically from the foregoing discussions.

B. Center the attention of the students on one of the reaction equations and ask them to propose possible mechanisms for the reaction between that metal and water. (The teacher should explain what is meant by the term mechanism --what happened to the reactants while they were becoming products.)

NOTE: The authors are not aware of any concrete experimental data that would support a specific mechanism for these reactions. The question and its discussion are proposed here for the purpose of involving the students in thinking about possible ways that the reactions could occur and produce the observed products. The discussion is valuable, also, in that it allows students to review previous learning and synthesize a considerable amount of information into a logical structure.

1. Encourage the students to analyze the reactants and products to determine what changes have occurred during the reaction.

2. Two possibilities for mechanisms would involve considering the hydrogen ion as the main reactant with the metal or, alternatively, considering molecular water as the main reactant.

3. This discussion should be allowed to continue until the teacher considers it to have been used to its maximum benefit.

C. The discussion guidelines for Steps 9 through 13 of the EXPLORATION are as follows:

1. Begin the discussion by performing a demonstration of the reaction of lithium with HCl and lithium with water. (These can be performed simultaneously to facilitate direct rate comparison.) This demonstration relates to the question asked in Step 6. Students will probably have predicted that metals generally react easier in acid than in water. If so, this demonstration will reinforce their generalization.

2. Ask the class to explain how (if at all) the increased speed of reaction between acid and lithium supports mechanisms that were proposed for the I-A metal reactions with water.

3. Ask if HCl was the only reactant with lithium in this demonstration. The idea here is to encourage the students to account as fully as possible for the chemical activity in the system. This situation can be related to previous experiments (for example, the reaction of Mg with air, the $CuSO_4$ catalysis of the reaction of zinc with hydrochloric acid and the reaction of calcium metal with hydrochloric acid solution).

4. Ask the students to review orally what happens to the valence electrons of the II-A and III-A metals when they react. Ask if aluminum chloride could be covalent. Continue the questioning to ascertain its polarity and consequent solubility characteristics.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

NOTE: A firm decision as to whether aluminum chloride is ionic or covalent is neither necessary nor desirable. It is interesting to note that either result can be achieved logically and with reference to experimental information. The discussion suggestions stated here for aluminum chloride are only a few of many possibilities. There are many idea expansions that can be drawn out of this type of discussion.

5. Have the class go over factors that will affect the ease with which electrons can be removed from these metal atoms.

6. Guide the class, through questioning and discussion, to the establishment of the general tends in reactivity for the metals in the I-A, II-a, and III-A columns (that is, reactivity increases as one moves down and to the left in the periodic table).

INVESTIGATION 19: PERIODICITY AND REACTIVITY

EXPANSION OF THE IDEA

A. Assign Reading a, REACTIVITY OF METALS. Discuss the reading with respect to the present investigation.

B. Experiment A: Halogens and Halides

INVESTIGATION 19: PERIODICITY AND REACTIVITY

Name _____

Section _____

EXPANSION OF THE IDEA

HALOGENS AND HALIDES: EXPERIMENT A

1. Test the solubility of NaCl, NaBr, and NaI in water by putting a small amount (enough to cover a match head) of each into a test tube and adding 2-3 ml of water and mixing vigorously. Repeat the test with 2-3 ml of hexane. Record your observations in the table below.

	water	hexane
NaCl	_____	_____
NaBr	_____	_____
NaI	_____	_____

2. Mix the NaCl in water and NaCl in hexane together. Mix the solution vigorously for 15 seconds. Do the same for the NaBr and NaI solutions. Record your observations.

3. Test the solubility of I_2 crystals in water and in hexane in a manner similar to Step 1. Add a few ml of hexane to the water solution of I_2 and mix vigorously. Record your observations.

BEST COPY AVAILABLE

INVESTIGATION 19: PERIODICITY AND REACTIVITY

INVESTIGATION 19: PERIODICITY AND REACTIVITY

4. Obtain water solutions of I_2, Br_2 and Cl_2. Place 5 drops of each in separate 13 x 100 mm test tubes and add 1 ml of hexane. Mix vigorously. Repeat the test using 1 ml of I_2, Br_2 and Cl_2 instead of 5 drops. Record your observations in the table below.

NOTE: These solutions should be retained as standards for comparison throughout the experiment. The test tube containing the color standards should, therefore, remain stoppered. If the color in the hexane layer should fade, add another 1 ml of the appropriate halogen water solution and shake.

	5 drops	1 ml
I_2:	_____	_____
Br_2:	_____	_____
Cl_2:	_____	_____

QUESTIONS

1. Classify, in the table below, water, hexane, fluorine (F_2), chlorine (Cl_2), bromine (Br_2), iodine (I_2), sodium fluoride (NaF), sodium chloride (NaCl), sodium bromide (NaBr), and sodium iodide (NaI), by listing them in the appropriate column.

Polar	Non-Polar

247

INVESTIGATION 19: PERIODICITY AND REACTIVITY

2. Use the test tube diagram to illustrate what you would see if equal volumes of hexane and water were poured into a test tube. Label your diagram.

3. Diagram and label what you would expect to find in the test tube of Question 2 if you added Br_2 to the tube or if you added NaBr to the tube.

4. Explain the chemical difference between the chlorine atom of NaCl and the chlorine atom in Cl_2.

248

NOTE: Although these questions should not be overly challenging to most students, the information contained in the answers to these questions will be quite important during the performance of the experiments with the halogens. Encourage the students to apply this information, as appropriate, while doing the laboratory work.

1.

Polar	Non-Polar
water	hexane
NaF	fluorine
NaCl	chlorine
NaBr	bromine
NaI	iodine

2. Hexane is insoluble in water. It is, furthermore, less dense than water, so the diagram should show a layer of hexane over a layer of water.

3. The VII-A elements, in their elemental states, would be soluble in hexane and the halogens are non-polar. According to the solubility rules, water would not be expected to be a good solvent for the halogens.

4. The chlorine atom in NaCl has picked up an electron to become a negative ion. The chlorine atoms in Cl_2 are each sharing one electron from the other atom, so that they could both be considered as being neutral.

NOTE: It is very important that the teacher carefully observe the students' progress with the laboratory work. Ask leading questions that will aid the students in relating observational information to chemical activity in the systems.

BEST COPY AVAILABLE

C. Experiment B

INVESTIGATION 19: PERIODICITY AND REACTIVITY

HALOGENS AND HALIDES: EXPERIMENT B

You will first need to make color standards to use for comparison with the reaction test tubes. According to the solubility rules, non-polar substances should dissolve in non-polar substances, so the color obtained in the hexane layer of these color standards should represent the color of the respective elemental halogens. Keep this fact in mind while performing the tests later.

Color Standards

1. Clean and dry three 13 x 100 mm test tubes. Label one test tube chlorine, a second test tube bromine, and a third test tube iodine.
2. Put approximately 1 ml of hexane in each of the test tubes.
3. To the chlorine test tube add approximately 2 ml of chlorine water; to the bromine test tube add approximately 2 ml of bromine water; and to the iodine test tube add approximately 2 ml of iodine water.
4. Stopper each test tube and shake vigorously.

NOTE: These standards should be retained for comparison throughout the experiment. The test tubes containing the color standards should, therefore, remain stoppered. If the color in the hexane layer should fade, add another 1 ml of the appropriate halogen water solution and shake.

5. Put approximately 1 ml of NaF solution in a clean, dry 13x100 mm test tube, 1 ml of NaBr solution in a second test tube, and 1 ml of NaI solution in a third test tube. To each of these test tubes add approximately 1 ml of hexane. Stopper, shake vigorously, and record your observations.

269

6. Add approximately 1 ml of Cl_2 water to each test tube. Stopper and shake vigorously. Record your observations of the systems after shaking.
7. In which test tubes did chemical changes occur? Support your answers with arguments based on your observations.
8. Write balanced equations for the chemical reactions. Include only the reactants and products in the equations, not the solution media.

NOTE: Dispose of the contents of these test tubes according to your teacher's instructions. Wash the test tubes.

270

9. Obtain approximately 1 ml of NaCl and NaI solutions in separate test tubes. Add 1 ml of hexane to each test tube, stopper, shake vigorously, and record your observations.
10. Add 1 ml of Br_2 water to each test tube, stopper, shake vigorously, and record your observations after shaking.
11. In which test tubes did chemical reactions occur? Support your answers with arguments based on your observations.
12. Write balanced equations for the chemical reactions. Include only the reactants and products, not the solution media.

NOTE: Dispose of the contents of these test tubes according to your teacher's instructions. Wash the test tubes.

271

6. Tell the students to keep the lids screwed on the dropper bottles except when they are in actual use. Elicit an explanation for this necessity (this can be explained in terms of the weak attractions holding the halogens in solution and the volatility of the elemental halogens).

7. When the students come in for the second day of laboratory work on this activity, lead them through the significance of the observations on one of the reactions they have already performed. The teacher may wish to carry this discussion to the extent of writing the reactants and products on the board. If this much help is given, however, the balancing of the equation should be left to the students.

   Reactions took place in the bromine and iodine test tubes. Chlorine water was added to these test tubes, and, if no reaction had occurred, the hexane layer should have contained chlorine. In these test tubes the hexane layer contained bromine and iodine, respectively.

8. $Cl_2 + 2NaBr \longrightarrow 2NaCl + Br_2$

   $Cl_2 + 2NaI \longrightarrow 2NaCl + I_2$

11. Reaction occurred only in the iodine test tube. The reasoning parallels that of Step 7.

12. $Br_2 + 2NaI \longrightarrow 2NaBr + I_2$

BEST COPY AVAILABLE

BEST COPY AVAILABLE

13. Obtain approximately 1 ml of 6M NaF, NaCl and NaBr solutions in separate test tubes. Add 1 ml of hexane to each test tube, stopper, shake vigorously, and record your observations.

14. Add 1 ml of I_2 water to each test tube, stopper, shake vigorously, and record your observations after shaking.

15. In which test tubes did chemical reactions occur? Support your answers with arguments based on your observations.

16. Write balanced equations for the chemical reactions. Include only the reactants and products, not the solution media.

NOTE: Dispose of the contents of the test tubes according to your teacher's instructions. Wash out the test tubes containing the color standards and the reaction test tubes.

QUESTIONS

QUESTIONS

1. Write total electron configurations (not electron-dot diagrams) for neutral F, Cl, Br, and I atoms.

2. When halogen atoms react chemically to form ions, what changes occur in their electron configurations?

3. Rank the halogens (bromine, chlorine and iodine) in order from least to most reactive. Offer a thorough explanation for this trend in reactivity. (Laboratory observations will not be sufficient to completely answer this.)

4. Predict the reactivity of fluorine compared to the other members of the VIIA family. Offer reasons for your proposed behavior of fluorine.

BEST COPY AVAILABLE

15. No reaction in any of the test tubes.

16. No reactions.

1. F $1s^2 2s^2 2p^5$

   Cl $1s^2 2s^2 2p^6 3s^2 3p^5$

   Br(Ar) $3d^{10} 4s^2 4p^5$

   I(Kr) $4d^{10} 5s^2 5p^5$

2. They pick up another electron, forming ions with a negative one charge.

3. Iodine, bromine, and chlorine is the order from least reactive to most reactive. Halogens react by becoming negative. The observations show that, halide ions were forced to give up their electrons, thus becoming neutral. This is the opposite of the normal way of reacting, and in each case it was forced by another halogen. The electronegativity trend is increasing from iodine to chlorine, meaning that iodine would have the smallest attraction (of the three) for another atom's electrons. This would lead to the expectaion that iodine was less reactive than bromine and that bromine was less reactive than chlorine.

BEST COPY AVAILABLE

D. Class discussion of the laboratory activity can proceed according to the following guidelines.

1. Review the reaction systems and have the class orally review the reactions that occurred (including balanced equations, which should be written on the board). At this stage in the discussion, the teacher should rely heavily on observational evidence for chemical changes in these systems. Considerable discussion may be required before all students are able to correctly interpret these observations.

2. Discuss, through questioning, the electronic changes that occur when non-metals react to form ions. Have the class compare this (by review) to the electronic changes undergone by metals in their ion forming reactions.

3. Have the class orally review the factors that affected the reactivity of metals.

4. Center the attention of the class on the reaction equations that were written at the beginning of the discussion. Ask the students to determine which of the VII-A elements is most reactive. This discussion will require quite a bit of time, in all probability, and some sensitive interaction and astute questioning on the part of the teacher. A good method involves having the students propose mechanisms for one of the reactions. (This can follow a pattern analogous to the one used with I-A metals.)

5. When the most reactive halogen has been determined, the trend in reactivity for the halogens should be easily ascertained by combining experimental evidence and logic. This trend should be conrasted to the trend in reactivity for a family of A-column metals.

6. Establish, through discussion, comparative reactivities of various non-metal families. (A good example to use is the initial reaction of magnesium with air, in which both the nitride and the oxide ions formed, with the nitride being formed in much smaller quantity.)

7. Discuss the overall reactivity trends for respresentative metals and non-metals.

E. Assign Reading b, REACTIVITY OF NON-METALS.